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THE

ELEMENTS

OF

INORGANIC CHEMISTRY.

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THE

ELEMENTS

OF

INORGANIC CHEMISTRY.

BY

J. C. BUCKMASTER, F.C.S.

LATE STUDENT IN THE GOVERNMENT SCHOOL OF MINES:
CERTIFICATED TEACHER OF SCIENCE BY
THE DEPARTMENT OF SCIENCE AND ART: AND
EXAMINER IN CHEMISTRY AND PHYSICS IN THE
ROYAL COLLEGE OF PRECEPTORS.

LONDON
LONGMAN, BROWN, GREEN, LONGMANS, & ROBERTS.
1858

193. c. 9.

BODLEIANA BIBLIOTHECA

DEDICATION.

TO

THE REV. JAMES BOOTH, LL.D., F.R.S., &c.
THE VICARAGE, WANDSWORTH

AND

THE REV. THOMAS JACKSON, M.A.
PREBENDARY OF ST. PAUL'S; RECTOR OF STOKE NEWINGTON; AND
LATE PRINCIPAL OF THE TRAINING COLLEGE, BATTERSEA.

GENTLEMEN,

I DEDICATE this work to you in grateful remembrance of your past kindness.

J. C. BUCKMASTER.

AUGUST 1858.

PREFACE.

I HAVE prepared this little work in the hope that it may be found useful in promoting the scientific education of the senior classes in trade schools. I trust it will not be without its value to pupil teachers, school-masters, and the students of classes in Mechanics' Institutions.

Having been engaged for the past seven years as a teacher of physics, both to boys and adults, I have always found the experimental sciences, especially chemistry, to be very popular subjects of instruction; and if this little book should contribute, in however small a degree, towards the scientific industrial education of the working classes, or stir up in some obscure worker a desire to learn, I shall be content to have done something for that class which has always had my warm and earnest sympathy.

For some years I have taken a deep interest in Mechanics' Institutions. I believe that many of them contain the elements of great usefulness, and may yet be made to take an important part in the secondary education of the working classes.

I recollect, when only fourteen years of age, attending courses of lectures at a Mechanics' Institution; and to these lectures I owe the desire to learn the little science I have since acquired.

The true elevation of the working classes chiefly depends on their own efforts, judiciously aided by other means. A large and liberal system of education would do much to remove many of those evils by which they are afflicted.

My strong faith in the education of the people has never once faltered.

The political privileges and liberty of the present day are quite equal to the social and intellectual condition of the people. To advance the one and neglect the other is to invert the natural order of society, and invest with power a people that have never been educated to use it.

I am not insensible to the objections which are frequently urged against the introduction of new subjects of instruction; but it has often struck me that some of the subjects usually taught in schools might be got rid of with great advantage to the children. The great object of education with the working classes is its practical value, and they rarely rise above this idea; nor do I think it so low and grovelling as some have described it. Unless you can instruct a boy in some of those principles and things which will help him through life, and supply him with new resources in the time of difficulty, which shall infuse a new sentiment into his labour, and relieve him from the monotony of mere mechanical manipulation,—unless this is done, education loses much of its usefulness, and fails in the estimation of those for whose benefit it is intended.

I would not, on this account, ignore every thing of an abstract character. A man who can demonstrate the first twenty propositions in Euclid, is raised to an intellectual level above that at which he stood before. But persons who have little capacity for the reasoning of mathematical investigation, often distinguish themselves in the sciences of experiment and observation; and these sciences are not without their logic in the hands of a good teacher. Working men are almost always experimentalists; a difficulty or failure is overcome by trying some other method; they seldom reason out a result,—indeed, results in natural science are rarely so obtained,—and if they reason at all, it is after the result is accomplished.

I have endeavoured to render this book as practical as possible; I am not aware that I have omitted anything of importance, and, in the hands of an intelligent teacher, it will admit of much amplification.

I must express my obligations to the lectures of Dr. Hoffmann for some of the information.

St. John's Hill, Wandsworth.

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ELEMENTARY CHEMISTRY.

THE earth is made up of a variety of solid, liquid, and gaseous materials. Chemistry investigates the relationship between the different kinds of atoms or particles of which these materials are composed. If the earth were composed of only one simple substance, there could be no such science as chemistry.

All substances, whether animal, vegetable, or mineral, are either simple or compound. Those which cannot be reduced into any simpler form are called elements, and these elements combine to form an infinite variety of compound bodies. We may illustrate this by the letters of the alphabet, which cannot be separated into any simpler form, but by different combinations form a great variety of *words*. These words can be separated into their letters, as can the compound bodies into the elements of which they are composed. It is only compound bodies which can thus be separated. Sulphur, iron, zinc, copper, oxygen, hydrogen, cannot be resolved into any simpler substance: they are therefore called elementary bodies. Water can be separated into oxygen and hydrogen, and water is therefore a compound body, and not an element, as was once supposed. Again, chalk, or carbonate of lime, is made up of three elementary bodies, — a metal called calcium, oxygen, and carbon. Common salt is composed of two elementary bodies,—a metal called sodium, and chlorine. The great majority of compounds are formed by the union of two, three, or four elementary bodies. The separation of these compound bodies into simpler forms is called analysis, and the formation of compound bodies by the union of simple ones is called synthesis.

The Elementary bodies are sixty-two in number; but many of them are so rare as not to require special attention. They are divided into metals and metalloids, or non-metallic bodies. The metals have great lustre: are good conductors of heat and electricity. These properties are never associated in a non-metallic body. This division is not based on any exact principles of science, but as a common means of description. The metalloids or non-metallic bodies in the following table are marked with an asterisk.

TABLE OF ELEMENTS.

Names of Elements.	Symbols.	Atomic Weights.	Specific Gravities.
Aluminum	Al.	13·7	2·60
Antimony (*stibium*)	Sb.	12·9	6·70
Arsenic	As.	75	5·88
Barium	Ba.	68·5	
Bismuth	Bi.	71	9·82
*Boron	B.	10·9	
*Bromine	Br.	80	2·97
Cadmium	Cd.	56	8·60
Calcium	Ca.	20	
*Carbon	C.	6	3·5
Cerium	Ce.	47	

Names of Elements.	Symbols.	Atomic Weights.	Specific Gravities.
*Chlorine	Cl.	35·5	2·47
Chromium	Cr.	26·7	5·9
Cobalt	Co.	29·5	8·54
Columbium (*tantalum*)	Ta.orCm.	184	
Copper (*cuprum*)	Cu.	31·7	8·89
Didymium	D.	49·6	
Erbium	E.		
*Fluorine	F.	19	
Glucinum	G.	26·5	3
Gold (*aurum*)	Au.	197	19·26
*Hydrogen	H.	1	0·069
*Iodine	I.	127	4·948
Iridium	Ir.	99	16
Iron (*ferrum*)	Fe.	28	7·8
Lanthanum	La.	47	
Lead (*plumbum*)	Pb.	103·7	11·35
Lithium	Li.	6·52	
Magnesium	Mg.	12·7	1·70
Manganese	Mn.	27·6	6·85
Mercury (*hydrargyrum*)	Hg.	100	13·59
Molybdenum	Mo.	46	8·6
Nickel	Ni.	29·6	8·28
Niobium	Nb.		
Nitrogen	N.	14	0·9729
Norium	No.		
Osmium	Os.	99·6	10
*Oxygen	O.	8	1·106
Palladium	Pd.	53·3	11·3 to 11·
Pelopium	Pe.		
*Phosphorus	P.	32	1·77
Platinum	Pt.	98·7	21
Potassium (*kalium*)	K.	39	0·865
Rhodium	R.	52·2	11
Ruthenium	Ru.	52·2	8·6
**Selenium*	Se.	39·5	4·28
*Silicon	Si.	21·3	
Silver (*argentum*)	Ag.	108	10·47
Sodium (*natrium*)	Na.	23	0·972

Names of Elements.	Symbols.	Atomic Weights.	Specific Gravities.
Strontium	Sr.	43·8	
*Sulphur	S.	16	2·07
Tellurium	Te.	64·2	6·25
Terbium	Tb.		
Thorium	Th.	59·6	
Tin (*stannum*)	Sn.	59	7·29
Titanium	Ti.	25	5·3
Tungsten (*Wolfram*) . .	Tn. or. W.	95	17·6
Uranium	U.	60	
Vanadium	V.	68·6	
Yttrium	Y.		
Zinc	Zn.	32·6	7·86 to 7·1
Zirconium	Zr.	22·4	

The rare elements are distinguished by italics, and the figures to the right represent the combining equivalent or atomic weights of these elements. The figures in the second column will be explained under the section on weight.

Some of the elements have a popular name: sulphur is called brimstone; mercury, quicksilver. Carbon when crystallised is a diamond; when prepared artificially, it is known as charcoal, coke, lamp-black, and bone-black. The Latin names of all the metals terminate in *um*. Five of the elementary bodies are gaseous under ordinary pressure and temperature;—hydrogen, oxygen, nitrogen, chlorine, and fluorine. Two are liquid,—mercury and bromine. The others are solid. The elementary bodies are represented by symbols. These symbols are of great importance in expressing chemical changes.

The student should learn the symbols and combining equivalents of the commonly occurring elements.

The elementary bodies forming a compound are frequently called its constituents, or components, and the act of union is called chemical combination, whilst the separation of a compound is called decomposition.

CHEMICAL AFFINITY.

The power or force by which bodies chemically unite is called affinity. It differs from the force of gravitation, which tends to draw all bodies towards the centre of the earth, and operates at all known distances. It also differs from cohesion, which binds like particles of matter together. The rain falls by gravitation, the particles of iron are held together by cohesion; but in iron rust (or oxide of iron) the oxygen is united to the iron by the force of chemical affinity. Chemical combination may be distinguished from mere mechanical mixture by the formation of compounds having properties different from the bodies which have entered into combination. And what is most remarkable, bodies of opposite chemical characters evince the strongest tendency to combination. Chemical affinity may be regarded as that force by which bodies unite, and form compounds having properties totally different to the constituents which have entered into combination. This combination is usually attended with an elevation of temperature. Chemical affinity acts with different degrees of force. If sulphuric acid be united with ammonia, and magnesia added to the compound, the ammonia is set free, and the sulphuric acid combines with the magnesia. This union of magnesia and sulphuric acid may be again decomposed with lime; and this compound may be decomposed with soda, then potash, then strontia, then baryta. From this it appears that the affinity of sulphuric acid for baryta is strongest, and for ammonia the weakest.

The force of chemical affinity is very much influenced by heat. Some chemical compounds are entirely decomposed by heat; the particles seem to be removed beyond the sphere of chemical attraction by the repulsive force of heat. Heat destroys all organic substances, but other compounds are formed of a more permanent character. Light, also, modifies the force of affinity: if equal volumes of chlorine and hydrogen be mixed in the dark they do not unite, but under the influence of sun-light combination takes place with great violence. The particles of matter attracted are assumed to be in different electrical states, and this is regarded as the basis of all chemical action. Electrical currents are among some of the most powerful causes of chemical decomposition. As chemical combination can only take place when the particles or molecules are brought within the sphere of chemical attraction, which acts at insensible distances, the force of cohesion must, first, be overcome; and this is done by dissolving the substance usually in water: so that the particles may be brought into closer contact. Substances in a solid state rarely combine, even if reduced to powder. Phosphorus with Iodine is an exception. When elements are liberated at the same time, in what is called their nascent condition, compounds are sometimes formed which cannot be effected under other circumstances. If nitrogen and hydrogen be mixed in the same vessel, combination does not take place; but when these gases are set free at the same time, as in the case of manure heaps, ammonia is formed.

CATALYSIS.

Chemical combination is also influenced by what is called catalysis; that is, a body by its mere presence induces changes, in which the body itself seems to act by mere contact. If starch be boiled in a little weak sulphuric acid, it is converted into sugar. This acid, if examined after the process has terminated, is found unaltered; so that the smallest quantity of acid is sufficient to convert any quantity of starch into sugar. The acid seems to act by its presence. The binoxide of manganese undergoes no change when heated with the chlorate of potash to obtain oxygen, but the oxygen comes off at a much lower temperature than when the salt is used alone.

LAWS OF CHEMICAL COMBINATION.

Chemical combination takes place according to certain definite laws.. These laws are called the laws of combining proportion: no circumstances can alter the proportional quantities in which bodies chemically unite. The same chemical compound invariably consists of the same elements united in the same proportion by weight. Pure water, no matter where obtained, contains 8 parts of oxygen by weight united to 1 part of hydrogen. Common salt, whether from Poland or Cheshire, contains 23 parts by weight of sodium, to 35·5 parts by weight of chlorine. Carbonate of lime, whether crystallised or amorphous, has the same elements united in the same proportion by weight.

When one body is capable of uniting with another, in several proportions, these proportions bear a simple relation to each other. This is called the law of multiple proportions. The best illustration of this law are the five oxygen compounds of nitrogen.

$$\underbrace{N = 14 \quad . \quad O = 8 \quad NO = 14 + 8 = 22}_{\text{Protoxide of nitrogen.}}$$

$$\underbrace{N = 14 \quad . \quad O_2 = 16 \quad NO_2 = 14 + 8 + 8 = 30}_{\text{Binoxide of nitrogen.}}$$

$$\underbrace{N = 14 \quad . \quad O_3 = 24 \quad NO_3 = 14 + 8 + 8 + 8 = 38}_{\text{Nitrous acid.}}$$

$$\underbrace{N = 14 \quad . \quad O_4 = 32 \quad NO_4 = 14 + 8 + 8 + 8 + 8 = 48}_{\text{Hyponitric acid.}}$$

$$\underbrace{N = 14 \quad . \quad O_5 = 40 \quad NO_5 = 14 + 8 + 8 + 8 + 8 + 8 = 54}_{\text{Nitric acid.}}$$

In this example, while the nitrogen remains constant, the oxygen increases by multiples of eight.

Sometimes the members of a group are wanting: if these members should be discovered, they will, doubtless, follow the law just illustrated. Chlorine affords a good example.

$$\underbrace{Cl = 35{\cdot}5 \ . \ O = 8 \quad ClO = 35{\cdot}5 + 8 = 43{\cdot}5}_{\text{Hypochlorous acid.}}$$

$$\underbrace{Cl = 35{\cdot}5 \ . \ O_3 = 24 \quad ClO_3 = 35{\cdot}5 + 8 + 8 + 8 = 59{\cdot}5}_{\text{Chlorous acid.}}$$

$$\underbrace{Cl = 35{\cdot}5 \ . \ O_4 = 32 \quad ClO_4 = 35{\cdot}5 + 8 + 8 + 8 + 8 = 67{\cdot}5}_{\text{Hypochloric acid.}}$$

$$\underbrace{Cl = 35{\cdot}5 \ . \ O_5 = 40 \quad ClO_5 = 35{\cdot}5 + 8 + 8 + 8 + 8 + 8 = 75{\cdot}5}_{\text{Chloric acid.}}$$

$$\underbrace{Cl = 35{\cdot}5 \ . \ O_7 = 56 \quad ClO_7 = 35{\cdot}5 + 8 + 8 + 8 + 8 + 8 + 8 + 8 = 91{\cdot}5}_{\text{Perchloric acid.}}$$

Here the compounds between the first and second, and the fourth and fifth are wanting. While the chlorine is a constant quantity, the oxygen increases by multiples of eight.

The oxygen compounds of sulphur afford another illustration of this law.

$$\underbrace{S_2 = 32 \quad . \quad O_2 = 16 \quad S_2O_2 = 16 + 16 + 8 + 8 = 48}_{\text{Hyposulphurous acid.}}$$

$$\underbrace{S = 16 \quad . \quad O_2 = 16 \quad SO_2 = 16 + 8 + 8 = 32}_{\text{Sulphurous acid.}}$$

$$\underbrace{S_2 = 32 \quad . \quad O_5 = 40 \quad S_2O_5 = 16 + 16 + 8 + 8 + 8 + 8 + 8 = 72}_{\text{Hyposulphuric acid.}}$$

$$\underbrace{S = 16 \quad . \quad O_3 = 24 \quad SO_3 = 16 + 8 + 8 + 8 = 40}_{\text{Sulphuric acid.}}$$

If one body, A, unite with other bodies, B C D, the quantities of B C D which unite with A represent the proportions in which they unite among themselves, in the event of union taking place.

No element has a greater range of affinity than oxygen, and many of its compounds are capable of exact analysis. Its combining equivalent is eight.

Let us take a few illustrations.

O = 8
H = 1. . . . S = 16
I = 127. . . . K = 39
P = 32. . . . Fe = 28
Cl = 35·5 . . . Ag = 108
N = 14. . . . C = 6

These numbers represent the proportions in which these elements unite: thus, 8 of oxygen unites with

1 of hydrogen, with 14 of nitrogen, with 6 of carbon, with 16 of sulphur, 39 of potassium, 28 of iron, and 32 of phosphorus. 1 equivalent of hydrogen goes as far in saturating 8 of oxygen, as 108 of silver, 39 of potassum, 6 of carbon, 32 of phosphorus, and 28 of iron. 16 of sulphur combines with 28 of iron, 35·5 of chlorine. This law is never departed from in any case. It is very common to speak of such a number of equivalents of one body united to one or more equivalents of another. Sulphuric acid contains one equivalent of sulphur to three equivalents of oxygen, SO_3. Potash, KO, one of potassium to one of oxygen. Phosphoric acid, PO_5, one of phosphorus to five of oxygen.

The combining equivalent of a compound is always the sum of the combining equivalents of its constituents. The combining equivalent of sulphuric acid is 40, 16 of sulphur, and 3 times 8 of oxygen. $S\ 16 + 8 + 8 + 8 = 40$. The combining equivalent of potash, KO, is 47, 39 the equivalent of potassium, and 8 the equivalent of oxygen, $39 + 8 = 47$. 40 parts of sulphuric acid neutralise 47 parts of potash, forming the sulphate of potash, which has a combining equivalent of $47 + 40 = 87$. Nitric acid, NO_5, has a combining equivalent of $14 + 8 + 8 + 8 + 8 + 8 = 54$; the proportion of nitric acid necessary to neutralise and combine with 47 parts of potash. From this it will be evident that when neutral salts decompose each other, the resulting compound will be neutral. It will be understood that these proportions are by weight.

COMBINATION BY VOLUME.

It is often more convenient to measure a certain volume of gas than to weigh it. Gases and volatile vapours combine by volume according to a definite and simple law. Equivalent weights of hydrogen and chlorine occupy at the same temperature and pressure equal volumes. Oxygen occupies only half the volume. If the specific gravity of a gas or vapour, and its combining equivalent, be known, its combining volume may be determined by dividing the chemical equivalent, of the substance by its specific gravity. The combining equivalent of oxygen is 8 divided by its specific gravity, $1{\cdot}106 = \frac{8{\cdot}000}{1{\cdot}106} = 7{\cdot}23$, its combining volume. Chlorine equivalent 35·5 divided by its sp. gr. which is $2{\cdot}470 = \frac{35{\cdot}50}{2{\cdot}470} = 14{\cdot}33$, its combining volume. Sulphur vapour, sp. gr. $\frac{16{\cdot}000}{6{\cdot}654} = 2{\cdot}40$.

From this it appears that the combining volume of chlorine is double that of oxygen; and sulphur one-sixth of chlorine. Hydrogen, nitrogen, chlorine, bromine, iodine, carbon, and mercury, in a gaseous state, have the same combining volume; oxygen, phosphorus, and arsenic, one-half; sulphur, one-sixth. Compound bodies in a state of vapour yield the same results.

WEIGHT.

By the term weight is meant the whole amount of matter a body contains, and this is generally expressed in grains. Specific weight, or specific gravity, is the number which expresses the ratio which the weight of a cubic inch of the body bears to the weight of a cubic inch of distilled water at a temperature of 60°.

The specific gravity of—

Platina	21
Gold	19
Lead	11
Copper	8
Iron	7

So that in the same bulk we have three times more matter in platina than iron, and iron is seven times heavier than its own bulk of water. In gases, air is taken as 1000 ; and the specific gravity of gases are referred to that standard. Great skill is required in determining the specific gravity of gases or vapours, and nothing short of practice can give accuracy in the results.

Atomic weight, or equivalent weight, is a term used to express the proportion in which bodies combine chemically. The atomic theory was first promulgated by Dr. Dalton. It assumes the existence of indivisible particles, molecules, or atoms ; that compounds are formed by the union of these atoms, one to one, one to two, one to three, and so on. The compound atom thus formed unites itself to the compound atom of another kind, and combination of the second order results. If it be allowed that the relative weights of these atoms are in the proportion of their equivalent numbers, the hypothesis renders satisfactory reasons for the laws of combining proportion. Chemical compounds must be definite in their composition, always containing the same number of atoms. The law of multiple proportion is also easily explained. An amto of nitrogen unites with one, two, three, four, and five of oxygen. If the black circle represent the nitrogen atom, and the open circles the oxygen, we may help our understanding by the following diagram:—

NO	Protoxide of Nitrogen	
NO_2	Binoxide of Nitrogen	
NO_3	Nitrous acid	
NO_4	Hyponitric acid	
NO_5	Nitric acid	

CHEMICAL NOMENCLATURE.

In early days, fanciful names were given to chemical compounds; many of these names have been retained in common use, such as oil of vitriol, aqua fortis, spirits of hartshorn, spirit of salt, cream of tartar, sugar of lead, blue stone, flower of zinc. Lavoisier introduced a new nomenclature, which satisfied for some time all the requirements of science. In organic chemistry, which is a new department of science, the defects of this nomenclature is sometimes felt. Some of the elementary substances have names referring to some peculiarity of the body: such as barium, antimony, chromium, chlorine, &c. When any of the non-metallic elements combine with a metal, the compound ends in *ide*. So we have oxides, bromides, chlorides, iodides, carbides, phosphides, and sulphides. Carburet, sulphuret, and phosphoret are sometimes used.

A compound of oxygen and iron, FeO, is the oxide of iron; by replacing the oxygen with chlorine, sulphur, or iodine, we get the chloride, sulphide, and iodide of iron. If the non-metallic bodies combine with each other, they are similarly named; as, sulphide of carbon, chloride of sulphur, phosphide of hydrogen.

The oxides are generally divided into three classes. The oxides which resemble potash, soda, and lime, are called basic oxides. The oxides of the second class are the very opposite in chemical relationship to the first, and are called acids. Phosphoric, sulphuric, chloric, and nitric acid may represent this class of oxides. There is a third class called neutral oxides, of which water and the binoxide of manganese are good illustrations. It frequently happens that a metal unites with oxygen in several proportions: in such cases, one oxide has generally a strongly marked basic character; to this the term protoxide is given. The compounds next are called binoxide or deutoxide, teroxide or tritoxide; the highest oxide, not having acid properties, is called the peroxide. Any compound containing less oxygen than the protoxide is called a suboxide. If an oxide should occur between the protoxide and binoxide, it is called the sesquioxide.

If the letter M represent a metal, and if oxygen unite with it in several proportions, we shall have —

M_2O suboxide, MO protoxide,
M_2O_3 sesquioxide, MO_2 binoxide,
MO_3 teroxide, MO_4 quadroxide

By substituting chlorine, or any of the other metalloids for the oxygen, we shall have the subchloride, protochloride, bichloride, &c.

When the metalloids unite, compounds are generally formed which have acid characters. The acid compound containing most oxygen has the termination in *ic*; as, sulphuric, nitric, chloric, phosphoric. The compound containing the next less quantity of oxygen terminates in *ous*; phosphorous, chlorous, sulphurous. When other compounds were discovered the prefixes *hyper* and *hypo* were introduced.

Hypophosphorous acid	PO
Phosphorous acid	PO_3
Phosphoric acid	PO_5
Sulphurous acid	SO_2
Sulphuric acid	SO_3
Hyposulphurous acid	S_2O_2
Hyposulphuric acid	S_2O_5
Protoxide of nitrogen	NO
Binoxide of nitrogen	NO_2
Nitrous acid	NO_3
Hyponitric acid	NO_4
Nitric acid	NO_5
Hypochlorous acid	ClO
Chlorous acid	ClO_3
Hypochloric acid	ClO_4
Chloric acid	ClO_5
Perchloric acid	ClO_7

The prefix *per* is usually given to indicate the highest or most intense. In the case of metals it is given to the highest oxide destitute of acid properties. The nomenclature of a salt is derived from that of the acid. When the acid ends in *ic* the salt ends in *ate*—sulphate of lime. When it terminates in *ous* the salt ends in *ite*—nitrite of soda. Hyponitrate, hyposulphate, hyponitrite, hypophosphite, &c.

Bisulphate, $MO.2SO_3$, sesquisulphate $2MO.3SO_3$.

BASIC OXIDES.

It has been stated that the metallic oxides are divided into bases, acids, and indifferent bodies. Several of the metals combine in different proportions with oxygen, and the chemical character of these compounds is determined in a great measure by the quantity of oxygen they contain. The oxides of manganese afford a good illustration. The protoxide, MnO, is a powerful base, the sesquioxide, Mn_2O_3, is feebly basic. The binoxide or peroxide, MnO_2, is an indifferent body; while MnO_3, manganic acid, and Mn_2O_7, permanganic acid, are acid bodies. Very few of the metals form acid oxides, or indifferent bodies; but the next oxide after an indifferent body, has acid properties. The most important basic oxides are the following. The symbols of these bases should be committed to memory.

List of Basic Substances.

Names.	Colour.	Symbols.
Potash	White	KO
Soda	Ditto	NaO
Ammonia	Ditto	NH_4O
Baryta	Ditto	BaO
Strontia	Ditto	SrO
Lime	Ditto	CaO
Magnesia	Ditto	MgO
Alumina	Ditto	Al_2O_3
Sesquioxide of chromium	Green	Cr_2O_3
Protoxide of iron	Black	FeO
Sesquioxide of iron	Brownish Red	Fe_2O_3
Oxide of zinc	White	ZnO
Protoxide of manganese	Greenish Grey	MnO

Names.	Colour.	Symbols.
Protoxide of nickel	Grey	NiO
Protoxide of cobalt	Ditto	CoO
Oxide of silver	Brown	AgO
Protoxide of mercury	Black	HgO
Binoxide of mercury	Red	HgO_2
Oxide of lead	Yellow or reddish yellow	PbO
Oxide of cadmium	Brown or yellowish brown	CdO
Oxide of copper	Black	CuO
Protoxide of bismuth	Yellow	BiO
Protoxide of tin	Black	SnO
Binoxide of tin	Light straw colour	SnO_2
Teroxide of antimony	Greyish white	SbO_3
Teroxide of gold	Brown	AuO_3
Binoxide of platinum	Brown	PtO_2

Most of these bases are insoluble in water: those which are soluble restore the blue colour of litmus, which has been reddened by an acid. Soda, potash, and ammonia are called alkalies: the latter is called the volatile alkali. These bodies are soluble in water, and neutralise the strongest acids. Baryta, strontia, lime, and magnesia, are called alkaline earths: except magnesia, they are all soluble in water, but less so than the alkalies. Alumina is insoluble in water, and belongs to the earths.

It may appear strange that the terms protoxide of potassium, protoxide of sodium, protoxide of ammonium, are not applied instead of potash, soda, ammonia, baryta and lime. There can be no doubt but protoxide of potassium is the correct term for potash, but when these names were given they were considered simple bodies, and the terms potash, lime, still continue in use. In writing out simple chemical formula, figures frequently occur. A little figure to the right merely shows the number of equivalents of the element taken, which stands to the left of the figure. Carbonic oxide, CO, means one equivalent of carbon and one of oxygen. Carbonic acid, CO_2, means one equivalent of carbon and two of oxygen. Sesquioxide of iron, Fe_2O_3, means two equivalents of iron and three of oxygen. The binoxide of mercury, HgO_2, means one equivalent of mercury and two of oxygen. The biniodide of mercury, HgI_2, means one equivalent of mercury and two of iodine. The terchloride of gold, $AuCl_3$, means one equivalent of gold and three of chlorine. The terhydride of arsenic, AsH_3, means one equivalent of arsenic and three of hydrogen. Hydrides must not be confounded with hydrates. The former is a union of hydrogen with some other body; the latter is a union of water with some substance, as the hydrate of lime, CaO,HO. When compounds of the metalloid and metal are written symbolically, the metal is placed first. Chloride of calcium, $CaCl$.

Exercise.

Write out the names of these compounds:—

CaS.
Fe_2Br_3.
AgI.
$SnCl_2$.
$PbCl$.
NH_4Cl.
CaF.
$SnCl$.
NH_4S.
CuO_2.
$AuBr_3$.
Cr_2O_3
FeS.
Mn_2O_3.
Fe_2S_3.
CaF

When the metalloids combine, the hydrogen is placed before the other metalloid, except in the case of phosphorus and carbon, CH. PH.

ACID BODIES.

An acid body was formerly defined as an oxidised compound which had a sour taste, turned blue litmus red, and neutralised alkalies. This definition is not sufficient: because we know many compounds without oxygen that have all the properties of an hydrochloric acid, HCl, and some bodies, such as carbonic and boracic acid, cannot neutralise the alkaline reaction of potash. Chemists now regard as acids all those compounds which unite with potash and soda, and give rise to compounds similar in their composition and character to the salts which sulphuric, nitric, or some admitted acid forms when combined with these alkalies.

The oxygen combinations with the metalloids form acids with very few exceptions. The hydrogen compounds of the metalloids are also chiefly acid. When oxygen acids combine with water, they are called hydrated acids, hydrated sulphuric acid, HO,SO_3. When they are deprived of this water, they are anhydrous, SO_3. Acids are monobasic, bibasic, or tribasic, according as they unite with one, two, or three equivalents of base. All the acids in the following list are monobasic, except phosphoric, arsenic, and arsenious acid.

OXYGEN ACIDS.

Sulphurous acid	gas		SO_2
Sulphuric acid [oil of vitriol]	fluid		HO,SO_3
Nitric acid [aqua fortis]	fluid		HO,NO_5
Chloric acid	fluid		HO,ClO_5
Oxalic acid	solid		$HO,C_2O_3=HO,\overline{O}$
Phosphoric acid	solid	tribasic	$3HO,PO_5$
Carbonic acid [choke damp]	gas		CO_2
Arsenious acid	solid	bibasic	$2HO,AsO_3$
Arsenic acid	solid	tribasic	$3HO,AsO_5$
Chromic acid	solid		CrO_3
Boracic acid	solid		BO_3
Silicic acid [silica, quartz, sand]	solid		SiO_3

Hydrogen Acids.

Hydrochloric acid [muriatic acid, spirit of salts]	gas	HCl
Hydrosulphuric acid [sulphuretted hydrogen]	gas	HS
Hydrofluoric acid	fluid	HF
Hydriodic acid	gas	HI
Hydrobromic acid	gas	HBr
Hydrocyanic acid [prussic acid]	fluid	$HC_2N = HCy$

SALTS.

There are two classes of salts, oxygen and haloid salts. Oxygen salts are formed by the union of an oxygen acid with an oxygen base, KO, SO_3, sulphate of potash. Haloid salts are formed by the radical of a hydrogen acid united to a metal, $NaCl$, chloride of sodium: here the radical of HCl, hydrochloric acid, is chlorine, which united to sodium forms common salt,—the type of this class of salts. Nearly all salts are solid at the common temperature, and have different affinities for water. The nitrate of lime and the nitrate of magnesia, the chloride of calcium and the silicate of soda, attract sufficient moisture from the atmosphere to become liquid: others undergo scarcely any change. Salts differ greatly in their degrees of solubility; some dissolve in less than their weight of water, while others are quite insoluble. Soluble salts crystallise out when their solutions are evaporated. The form which they assume in crystallising is always constant, and forms a character by which they are frequently distinguished. Some salts, in crystallising, unite with a definite quantity of water, called water of crystallisation. This water of crystallisation is held by a very feeble affinity: a moderate heat is sufficient to expel it. A portion of water is frequently retained with such force that it cannot be expelled without decomposing the salt. When this is the case, it is called essential water of crystallisation, and is expressed by the symbol HO. When the water is non-essential, it is expressed by the Latin word *aqua*. The following formulæ of salts must be studied. The letter M stands for a metal.

General Formulæ of the more important Salts.

OXYGEN SALTS.

Sulphites	MO,SO_2	Carbonates	MO,CO_2
Sulphates	MO,SO_3	Arsenites	$2MO,AsO_3$
Nitrates	MO,NO_5	Arseniates	$3MO,AsO_5$
Chlorates	MO,ClO_5	Chromates	$MOCrO_3$
Oxalates	$MO,\overline{O}$	Borates	MO,BO_3
Phosphates	$3MO,PO_5$	Silicates	MO,SiO_3

HALOID SALTS.

Chlorides (muriates)	MCl	Iodides	MI
Sulphides (sulphurets)	MS	Bromides	MBr
Fluorides	MF	Cyanides	MCy

It is seen by these formulæ how the terminations *ic* and *ous* in the acids change into *ite* and *ate* in salts. Oxygen acids unite with metallic oxides, and form oxygen salts. Write out the following salts:—

NaO,SO_2.
NH_4O,SO_3.
BaO,SO_3.
CaO,SO_3.
MgO,SO_3.
ZnO,SO_3.
FeO,SO_3.
CuO,SO_3.

KO,SO_2.
NaO,SO_2.
NH_4O,SO_2.
AgO,CO_2.
SrO,ClO_5.
HgI.
HgO,CrO_3.
KO,ClO_5.

KO,NO_5.
ZnS.
HgS_2.
$FeCl$.
$NH_4O,\overline{O}$.
$NoO,\overline{O}$.
KCy.
$AnCy_3$.

$HgCy$.
Fe_2Cy_3.
NiO,SO_3.
CdO,SO_3.
$HgCy_2$.
PbI.
$PbO,\overline{O}$.
BaO,NO_5.

There are three kinds of salts; neutral salts, acid salts, and basic salts. All salts are now regarded as neutral when they contain as many equivalents of acid as there are equivalents of oxygen in the base, $KOSO_3$, $Al_2O_33SO_3$, SnO_22SO_3. If no prefix is employed to indicate the quantity of acid or base, a neutral salt is always meant: thus, protosulphate of iron, FeO,SO_3; sulphate of the sesquioxide of iron, $Fe_2O_3,3SO_3$.

Exercises.

Write out sulphate of alumina, sulphate of binoxide of mercury, sulphate of teroxide of antimony, chlorate of protoxide of mercury, nitrate of the sesquioxide of iron, sulphate of the sesquioxide of chromium. An acid salt contains more equivalents of acid than there are equivalents of oxygen in the base.

There are two kinds of acid salts, one in which the dry acid is in excess, the other in which it is in excess in the hydrated state KO,CO_2,HO,CO_2; the bicarbonate of potash in which the hydrated acid is in excess; $KO,2CO_2$, the bicarbonate of potash, in which the dry acid is in excess. The hydrated acid salts are now regarded as double salts, in which one base is water. The prefix *bis* is employed when the salt contains two equivalents of acid to one of base, as bisulphate of soda, $NaO,2SO_3$. Biborate of magnesia, $MgO,2BO_3$. *Sesqui* is employed to indicate two of base to three of acid, as sesquicarbonate of ammonia, $2NH_4O,3CO_2$. The prefix *ter* is used to indicate three of acid to one of base, as terborate of magnesia, $MgO,3BO_3$.

Basic salts contain an excess of base. The terms bibasic and tribasic are used to indicate this excess;

as, bibasic phosphate of lime, $2CaO,PO_5$, tribasic arseniate of soda, $3NaO,AsO_5$. All salts are basic which contain more than one atom of base for one of acid : we have bibasic arseniate of lime, $2CaO,AsO_5$, bibasic arseniate of soda, $2NaO,AsO_5$, tribasic phosphate of soda and ammonia, NaO,NH_4O,HO,PO_5. In polybasic salts the base may be, as we have just seen, of different kinds. There are double oxygen salts, and double haloid salts. Sulphate of soda and lime, $NaO,SO_3 + CaO,SO_3$. Compounds of two elements are called binary compounds, as all the oxides and haloid salts, potash, KO, chloride of sodium, NaCl. When two binary compounds unite, a ternary compound is formed, sulphate of potash, KO,SO_3 oxygen salts. A quaternary compound is formed when two ternary compounds combine together as all double oxygen salts, $Al_2O_3,3SO_3+KO,SO_3$.

The plus sign is used in formulæ to show the arrangement of the elements and their degree of affinity in the compound. The comma is used to separate ternary compounds into binary, and to indicate the force of affinity which binds the elements or compounds together, KO,NO_5. The potassium and oxygen are united by a greater affinity than the nitric acid and potash, and it would be much easier to separate the salt into nitric acid and potash, than into nitrogen, oxygen, and potassium. When a large figure is placed on the line it multiplies all the elements on the right of it as far as the first comma or plus sign ; if before a bracket it multiplies all in that bracket, $2SO_3$, two equivalents of anhydrous sulphuric acid; $3CO_2$, three equivalents of carbonic acid ; $3FeO,CO_2$, means three equivalents of protoxide of iron and one of carbonic acid ; but $3(FeO,CO_2)$ signifies three equivalents of the protoxide of iron, and three equivalents of carbonic acid — in short, three equivalents of the proto-carbonate of iron. When more than one equivalent of any ternary compound is expressed, it should always be placed in brackets with the figure outside:—

$3AuCl_3$.—Three equivalents of the terchloride of gold.

$2Cr_2Cl_3$.—Two equivalents of the sesquichloride of chromium.

$2(PtCl,KCl)$.—Two equivalents of the chloride of platinum and potassium.

$2FeI$.—Two equivalents of the proto-iodide of iron.

$3(NH_4O,NO_5)$.—Three equivalents of the nitrate of ammonia.

$3(Fe_2O_3,3\bar{O})$ or $3(Fe_2O_3,3C_2O_3 + HO)$. — Three equivalents of the oxalate of the sesquioxide of iron.

$2MgO,3SiO_3+2HO$. — One equivalent of the sesquisilicate of magnesia, and two of water.

$3(Al_2O_3,3SO_3 + KO,SO_3 + 24aq)$.—Three equivalents of the sulphate of alumina and potash and 24 equivalents of non-essential water.

Interpret the following formulæ :—

$3AsH_3$.
BaO,CO_2+CaO,CO_2.
Cu_2I.
NaO,HO.
$3(MgO,CO_2)+aq$.
KO,C_2O_3+HO, or KO,O.
$CaO,3SiO_3$.
$2CuO,HO,PO_5$.
$2MgO,NH_4O,PO_5$.
MnO,SO_3+NaO,SO_3.
$Cr_2O_3,3HO$.
$3(2CuO,AsO_3)$.
$2NaO,HO,PO_5$.
CaO,C_2O_3+2HO.
CaO,CO_2+3MgO,CO_2.
$Al_2O_3,3HO$.

HEAT.

The term Caloric was introduced by Lavoisier to distinguish the cause of heat from the sensation. The terms caloric and heat are now used as synonymous.

The greater number of bodies, gaseous, liquid, and solid, expand when heated. Solids expand least, gases most; liquids seem to be intermediate.

Solids have an expansibility peculiar to themselves. Metals are the most expansible, zinc most, platinum least; glass, brick, porcelain, marble, stone, have small expansibilities. Gases undergo the same expansibility for additions of the same amount of heat, but liquids and solids have specific expansibility, that is, each has an expansibility peculiar to itself. Liquids, and solids, by the addition of the same heat, expand to a different extent. Solids and liquids expand more at high temperatures than at low. That is, if we had a piece of metal at the temperature of the air, and applied heat, it would expand to a less degree than if the heat had been applied when it was red hot.

The Thermometer.—A glass tube of uniform bore is taken, and one end blown into a bulb, in the blowpipe flame of a lamp. The bulb, when nicely formed, is carefully heated over the flame of a spirit lamp. This causes nearly all the air to be expelled. The open end of the tube is then placed in a vessel containing mercury. As the bulb cools the pressure of the air on the surface of the mercury drives a portion of the fluid into the bulb. By applying heat so as to cause the mercury to boil, the remainder of the air is driven out. Heat is now applied till the column of mercury rises to the top, it is then hermetically sealed by the blowpipe. The contraction of the mercury leaves a vacuum in the upper part of the tube which is essential to the perfection of the instrument. It has now to be graduated. Two points are taken. The temperature of melting ice or a mixture of ice and water;—the freezing point of water and the melting point of ice is always constant. This constitutes the freezing point. The boiling point of water is always constant, regard being had to the height of the barometer when great accuracy is required. This constitutes the boiling point. These points are marked, and the interval divided into degrees: the division is entirely arbitrary. In Europe and America the Centigrade scale is used; in England, the scale known as Fahrenheit's, is employed. In the F. thermometer the interval between the freezing and boiling points is divided into 180°, in the Centigrade it is divided into 100°. In F. the temperature of ebullition is expressed at 212°. In Russia, Reaumur's scale is used, which commences at 1° and the boiling point is at 80°. It is sometimes desirable to reduce these to the same scale.

$$\text{One degree of F.} = \tfrac{4}{9} \text{ of R.}$$
$$\text{ditto of F.} = \tfrac{5}{9} \text{ of the C.}$$

Multiply the number of degrees of F. above or below the freezing point by 4, and divide by 9; the sum will indicate degrees upon R.

$$\text{F. } 50 - 32 = 18 \times 4 = 72 \div 9 = 8^\circ \text{ R.}$$
$$\text{F. } 185 - 32 = 153 \times 4 = 612 \div 9 = 68^\circ \text{ R.}$$

$$\text{R. } 8 \times 9 = 72 \div 4 = 18 + 32 = 50^\circ \text{ F.}$$
$$\text{R. } 68 \times 9 = 612 \div 4 = 153 + 32 = 185^\circ \text{ F.}$$

$$\text{F. } 86 - 32 = 54 \times 5 = 270 \div 9 = 30^\circ \text{ C.}$$
$$\text{F. } 176 - 32 = 144 \times 5 = 720 \div 9 = 80^\circ \text{ C.}$$

$$\text{C. } 30 \times 9 = 270 \div 5 = 54 + 32 = 86^\circ \text{ F.}$$
$$\text{C. } 80 \times 9 = 720 \div 5 = 144 + 32 = 176^\circ \text{ F.}$$

A mercurial thermometer may be graduated up to 570° a few degrees above this point mercury begins to volatilise.

When very low temperatures have to be indicated, a spirit thermometer is used as the greatest attainable cold does not freeze the liquid, but it boils at 141°.

A thermometer reveals to us the intensity of heat, not the quantity, for a test tube of boiling water raises the mercury to the same point as a pail full, although the larger quantity must contain the most heat.

When directions are given to raise the temperature of a body to 110° it means that heat is added sufficient to raise the mercurial column to 110°.

Specific Heat.

Different substances require different degrees of heat to raise them to the same temperature. This is called specific heat.

If 1lb. of HO at 100° be mixed with 1lb. at 40°, the mean temperature will be $\frac{100^\circ + 40^\circ}{2} = 70^\circ$.

In the same way, the mean temperature of the same liquids may be found when mixed together, oil $\frac{80 + 20}{2} = 50^\circ$; mercury $\frac{120 + 40}{2} = 90^\circ$.

But if 1lb. of HO at 100° be mixed with 1lb. of olive oil at 40°, we shall not have the mean temperature of 70° but 80°. In the same way, if 1lb. of mercury at 40° be mixed with 1lb. of HO at 100°, we shall have a temperature of 98°. Now, the water parted with 20° of heat in the case of oil, which was sufficient to raise its temperature 40°, that is from 40° to 80°.

In the case of the mercury, the HO parted with only 2°, which was sufficient to raise the mercury 58° from 40° to 98° : equal weights of the liquids being taken in the experiment. From these experiments it appears that equal weights of HO, oil, and Hg, require different degrees of heat to raise them to the same temperature.

Solids may be easily experimented on if 1lb. of iron be taken and raised to a given temperature, and immersed in 1lb of HO at a known temperature, we can easily determine the specific heat by the increase of temperature, the same as by mixing two liquids at a different temperature. Specific heat is generally taken with reference to equal weights rather than equal volumes. 1 lb. of HO in rising to a given temperature, requires or absorbs thirty times more heat than 1 lb. of Hg. If Hg be taken as 1, water will be 30. The great specific heat of HO is very important: it rises in temperature slowly, and parts with its heat slowly, which tends to equalise the temperature of the air and earth. The small specific heat of Hg is useful. No relation exists in the table of specific heats when equal weights of different bodies are taken, but if *equivalent* weights be taken, then we have a relation. Let us take 16 oz. of S, 28 of Fe, 32 of P, 6 of C. Let these equivalent quantities be then raised to the temperature of 212°, and afterwards immersed in water at 60°, so as to observe how much each raises the temperature of the liquid. By such an experiment we shall find that 16 parts of sulphur and 28 of iron will raise the water to the same temperature, while the phosphorus will raise it four times higher. It thus appears that the specific heats of sulphur and iron, when taken in equivalent quantities, are the same, whilst phosphorus is four times greater.

Latent Heat.

When any solid body is converted into a liquid, a large amount of heat is observed to enter it without raising its temperature. This heat serves to liquefy the body without increasing its temperature. The water which flows from melting ice is no hotter than the ice; the heat it contains is not sensible, nor does it affect the thermometer; it is called latent heat. If a piece of ice be hung in a warm room it melts slowly. This would not be the case if a small amount of heat was only necessary to liquefy it. It gathers heat from surrounding objects, which is all expended in liquefying the body. If a cubic inch of ice be taken and a cubic inch of water, and the temperature of the water lowered to 32°, and the ice and water removed into a warm room, and a thermometer placed in the water, we shall be able to ascertain the amount of heat the water receives in a given time. Let us suppose the thermometer to rise 7° every half hour, and the time observed till the whole of the ice is melted. There will be 14° of heat added every hour: by these means we shall be able to ascertain how much heat disappears, or becomes latent in converting the ice into water. It will take 10½ hours to liquefy the ice. The water in this experiment was only 40°, or 8° above 32°. Now, if we multiply 7×21, we have $147 - 8 = 139$, the number of degrees of heat absorbed or rendered latent in converting the ice into water; in other words, 139° of heat have become latent. If we take 1 oz. of powdered ice or snow, and 1 oz. of water at 172°, the ice will be melted, and we shall have 2 oz. at 32°. The water lost 140° of heat, and the ice absorbed it; but if we take water at 32°, and water at 172°, we shall have the mean temperature of 102°. What we have said as regards ice, holds good with all solids: a certain amount of heat is in all cases rendered latent. The following table shows the latent heat of different bodies:—

Water	140
Sulphur	145
Lead	162
Wax	175
Tin	500

When water freezes, the 140° of heat abandon it, and manifest itself in a free or sensible state. If a flask of boiling water be saturated with the sulphate of soda, and the cork withdrawn and a rough body introduced, the soda will become solid, and the sudden conversion of the liquid into a solid causes it to part with its heat; if a thermometer be introduced the mercury in it will rise 30° or 40°; this is precisely what takes place in the freezing of water. Solids in becoming liquids render a large amount of heat latent, and liquids in becoming solids part with the heat previously rendered latent.

When solids are compelled to liquefy rapidly without a free supply of heat, the temperature of surrounding objects is lowered. Salt in water, or nitre, causes the thermometer to fall several degrees. If NaO,SO_3 and snow be mingled together the temperature falls to zero. $CaCl$ mixed with snow, lowers the temperature so as to freeze mercury. These are called freezing mixtures.

Latent Heat of Gases.—It is not necessary for a liquid to boil in order to produce vaporisation. The drying up of water by its gradual conversion into vapour is a phenomena with which we are all familiar.

If water be raised to 212° it is converted into steam, but the steam is the same temperature as the water. If it were only necessary to raise water to 212° to convert it into vapour, it would, as soon as it attained that temperature, explode into steam, like gunpowder. A cubic inch of water will make a cubic foot of steam. If vapours contained no more heat than the liquids, they would immediately condense into liquids as soon as they came in contact with any body lower in temperature; but it is found necessary to expose them to a great amount of cold. Ice in becoming water, renders a large amount of heat latent; water in becoming steam also renders a large amount of heat: what occurs with water occurs with all liquids.

If a vessel with water, at 32°, be placed over a regular source of heat, and the time observed necessary to raise it to 212°, it will require five times longer before the water is entirely converted into steam. Now the water has been raised through a temperature of 180°, say in one half hour, we multiply 180 × 5, which is the number of hours taken, and we get 900°: yet the steam has only the temperature of 212°. What has become of this enormous amount of heat? It has been rendered latent in the steam, and this can be demonstrated by condensing the steam, and observing the amount of heat evolved. Let the steam be conducted into ice-cold water, and it will part with its latent heat, and raise the water to its boiling point. If 11 cubic inches of water be taken at 32° and raised to 212° by the steam of 2 cubic inches of water, we shall have 13 cubic inches at 212°. 2 cubic inches of water, in the form of steam, have raised 11 cubic inches of water, mercury 180°, and the whole 13 cubic inches have a temperature of 212°. If a cubic foot of steam, at 212°, be contained in a close vessel, and 5½ inches of ice-cold water, at 32°, be injected into this vessel, we shall have 6½ cubic inches of water at 212°. Now it is evident that in returning to a state of water the steam has given out sufficient heat to raise 5½ cubic inches of water from 32° to 212°. The heat which was latent in the steam has now been made sensible in raising the 5½ cubic inches of water through 5½ times 180° of heat, viz. 990°.

CRYSTALLISATION.

In order that a body may crystallise it is necessary that the particles be free to move; when this occurs they arrange themselves in most cases in regular crystalline forms.

From this it appears that all crystalline bodies must at some time have been fluids. The regular formation of a crystal may be seen by splitting it, which may be easily done with some crystals. If the subdivision of the crystal be continued in those directions called planes of cleavage, a symmetrical figure is obtained which is the primitive form of the crystal, which may widely differ from the figure at first; should this be the case the latter will be the secondary form of the crystal. By subdividing the primitive form you get a number of minute fractions resembling it in figure; these are regarded as the type of the atoms of which the body is composed.

Among natural crystalline forms we may mention rock crystal, quartz, granite, calcareous spar, fluor spar, copper ore, pyrites, the loadstone, snow, and ice. Artificial crystals, — alum, common salt, saltpetre, sulphate of copper, sugar, sulphur, and bismuth, and a variety of others.

In the tinning of iron plates the crystallism of the tin is evident when acted on by nitric acid, producing the Mora antique, as also in the galvanising of iron. The planes of cleavage are seen in lead ore, Iceland spar.

When Iceland spar is broken it gives perfectly formed rhomboidal crystals.

The primitive form is in most cases taken as the basis of classification of different crystalline bodies. The chief points to be attended to in crystals are the angles and the centres of their edges and sides. The lines joining these points and passing through the middle of the crystal are called the axes. Most crystalline bodies have three such axes at right angles to each other.

These axes are either all equal, or two of them are equal, or all three unequal. That one most dissimilar from the other two is called the principal axis; the others secondary. In the cube either of the three may be regarded as the principal axis. A double pyramid has four.

The term amorphous is applied to bodies destitute of any regular form. Dr. Wollaston, however, proposed another theory; he regarded small spheres as the ultimate particles when placed in the same plane.

All metals under favourable circumstances assume a crystalline form.

Crystalline forms are generally arranged into six systems.

1. The regular system. The crystals of this division have three equal axes at right angles to each other; the cube, the regular octohedron, and the rhombic dodecahedron. Many substances, simple and compound, assume this form. The metals, also carbon, the diamond, common salt, iodide of potassium; the alums, fluor spar, bisulphide of iron.

2. The square prismatic system. These axes are all at right angles; two only are of equal length, peroxide of tin, the ferrocyanide of potassium

3. The right prismatic system. Three axes of unequal length; the right rhombic prism, — sulphur, nitrate of potash, sulphate of potash, sulphate of baryta.

4. The oblique prismatic system. Crystals belonging to this group, the oblique rectangular prism, the oblique rhombic prism — sulphate of soda, carbonate of soda, biborate of soda.

5. The oblique double prismatic system. Sulphate of copper.

6. The rhombohedral system. Ice, nitrate of soda, quartz, arsenic, antimony.

APPARATUS.

Expertness and neatness in the construction and use of apparatus can only be acquired by practice.

Corks are of constant use, and they should be of the best kind. Before fitting them into apparatus they should be softened by rolling them under the foot on a piece of paper, so as to keep them free from dust.

When glass tubes have to be passed through the cork, holes should be made with a cork-borer rather smaller than the tube to be inserted, and nicely eased to the proper size with a rat-tail file. If the cork should prove porous the exposed parts must be covered with a solution of sealing-wax in spirits of wine.

Glass tubing may be cut into the required lengths by marking round the tube with a three-cornered file; the tube is then easily broken. The depth of the mark made by the file must depend on the thickness of the glass. If the tube has to be bent this should be done before cutting it. To do this the tube must be heated gradually in the flame of a gas or spirit lamp, by turning it round in the flame and moving it a little to the right and left. As soon as the tube becomes red-hot it can be easily bent into the required shape. A little practice with a few pieces of glass tube over a lamp will soon give the experimenter more knowledge than any written instructions.

India-rubber tubing, which can be purchased any size, is used for connecting glass tubes. In connecting glass tubes by means of India-rubber the ends of the tube should not be brought so close together as to interfere with the flexibility of the arrangement. If the India-rubber tubing does not fit the glass tubes perfectly air-tight it must be tied with a bit of twine.

A good supply of Florence flasks will be found extremely useful, as they can be fitted up to suit a variety of purposes.

Gas Jars. — Plain cylindrical jars should be from twelve to sixteen inches in length and about two and a half to three inches in diameter, and about one-eighth of an inch in thickness. The open end should be ground, so as to be closed air-tight with a piece of flat glass. A little grease should be placed on the glass plate or edge of the gas jar, so as to ensure perfect tightness. Stoppered jars usually have glass stoppers like bottles, and should be ground at the top, so as to be closed with a glass plate in the manner described for cylindrical gas jars. These jars are removed from the pneumatic trough by passing a plate or saucer under the bottom. The plate or saucer should always contain a little water, so as to prevent the access of air and escape of the gas.

When substances have to be burned in a gas jar they are introduced by means of a deflagrating spoon. This spoon should be fitted with a plate, so as to cover the aperture of the jar and allow it to pass the required depth into the gas. In burning phosphorus a piece not larger than a split pea should be used; it should be carefully dried by gently wrapping it in blotting-paper. The phosphorus may then be ignited by touching it with a red-hot wire and quickly introducing it to about two-thirds the depth of the jar.

In collecting gases over water, where heat is required, the heat should in all cases be applied gradually, and the first portions of gas allowed to escape, till all the common air in the apparatus has been expelled.

The pneumatic trough should be about two feet long, eighteen inches wide, and a foot deep, provided with a shelf about three inches below the edge of the trough. The water should always stand about one inch over this shelf. Holes are made in this shelf, over which the jars intended to receive the gas are placed. The delivery-tube is brought under the open end of the jar, and as the gas rises in the jar the water is displaced. A spout should be provided to the trough, so as to allow the surplus water to pass into a pail. An ordinary gas jar is filled with air, and, as no two bodies can occupy the same space at the same time, it is necessary to expel this air before the jar can be filled with gas. This is done by immersing the jar sideways in the water of the pneumatic trough, turning it over with the open end downwards, and gently raising it to the shelf of the trough. If the stopper of the jar be air-tight the water, although above the level of that in the trough, will stand for any length of time in the jar. A little acquaintance with pneumatics will enable the pupils to understand the reason.

It is sometimes necessary to dry gases. This is effected by passing them through a series of Wolfe's bottles containing sulphuric acid, or passing the gas through a tube, from a foot to eighteen inches long and three-quarters of an inch in diameter, filled with the chloride of calcium. Tubes in the form of a U are frequently used. The sulphuric acid and the chloride of calcium are called desiccating agents, because they deprive the gases of water.

In the generation of carbonic oxide, from oxalic acid and hydrated sulphuric acid, carbonic acid, and carbonic oxide, are evolved simultaneously. The presence of the carbonic acid interferes with a careful examination of the carbonic oxide. This carbonic acid must be removed; and this is effected by passing the gases through a strong solution of caustic potash or lime water, which retains the carbonic acid. By this method pure carbonic oxide is obtained, which is passed into gas jars filled with water at the pneumatic trough.

When no heat is required to liberate gases or gaseous compounds, a two-necked Wolfe's bottle may be used. It should be fitted up with two glass tubes. The delivery-tube, from which the gas escapes, may be about a quarter of an inch in diameter, and pass just through the cork of the bottle; another tube, passing to within a quarter of an inch of the bottom of the bottle, and about half an inch in diameter, for pouring in acid or water. This saves the trouble of taking out the cork, which must be done every time fresh acid or water is added, unless the bottle be provided with two tubes in the manner described. The materials in the bottle should never occupy more than a third, and the acid should be added carefully, a little at a time, and the materials in the bottle agitated by a little shaking. A Florence flask may be fitted up with two tubes; but the neck is rather too small, unless great care be taken in boring the corks.

The whole of the apparatus for the evolution of gases must fit air-tight. Too much pains cannot be taken to secure this object; it will save much trouble and disappointment. The expression, "That's near enough," will not do in chemistry; it must be exactly, or a satisfactory result cannot be obtained.

Gases are sometimes collected by displacement.- If the gases are lighter than common air, as hydrogen and ammonia, or heavier, as chlorine and carbonic acid, the pneumatic trough may be dispensed with. But gases of nearly the same specific gravity of common air must be collected over water. The same generating apparatus as that just described will do, provided the flask or bottle will stand heat if required. The delivery-tube must pass to the bottom of the jar it is intended to fill with gas. This is often done by blocking the jar up with pieces of wood of different thicknesses. The top of the jar may be covered with a piece of card-board, with a hole for the delivery-tube to pass through. As the gas collects at the bottom it gradually expels all the air, exactly the same as if water were poured into the gas jar. In raising bottles or other apparatus let the blocking be firm, the largest and heaviest pieces at bottom. When the gases are lighter than common air the receiver may be supported by a retort-stand, the delivery-tube passing upwards to the extremity of the vessel intended to receive the gas. When large quantities of gas are required, a gasometer is used. The most convenient is the one known as Pepy's gas-holder. I shall not describe the action of this piece of apparatus, because five minutes' examination of one of these gasholders will give the student a much better idea of its action than any written description. The most careful explanation of any piece of mechanism or apparatus must full short of actual observation; and the teacher who attempts to teach without the objects, when they can be obtained, teaches words and not things.

In chemistry the terms solution, soluble, and solvent are frequently employed. Many solid bodies, when placed in water, dissolve. Salt and sugar dissolve in water. These substances are said to be soluble in water; and the water is called the solvent or menstruum. The more a solid is reduced to powder the easier it dissolves. If sugar be placed in alcohol, it does not dissolve. Here, then, we have an example, which may be multiplied to a great extent, of a body soluble in one liquid and insoluble in another. A solution may be either mechanical or chemical. The sugar in water is a case of mere mechanical solution; no chemical change has taken place, and, by evaporation, the sugar can be obtained in its original condition. Salt dissolved in water is another illustration of the same kind. In chemical solutions the solid and liquid combine together, forming a new compound, from which the original solid and fluid cannot be obtained by evaporation or any mechanical operation. When chalk is dissolved in hydrochloric acid, or copper dissolved with nitric acid, we have examples of chemical solution. A liquid can only dissolve a certain quantity of a solid substance. This quantity will vary according to the solid and liquid employed. When the solvent or liquid contains as much of the solid matter as it is capable of dissolving, it is said to be saturated, as all additional quantities of the solid fall to the bottom.

The solubility of most substances is increased by an increase of temperature.

100 pts.	of water at	60°F.	dissolve	11 pts. of sulphate of potash.
"	"	212°	"	25 " "
"	"	60°	"	32 pts. of dry sulphate of magnesia.
"	"	212°	"	74 " "

Common salt is about as soluble in hot water as cold. Some bodies attain their maximum solubility at a particular temperature. The solubility of the dry sulphate of soda is at its maximum at a temperature of 93°; but in the nitrate of soda we have the order reversed, — water, at 32°, dissolves 80 parts; at 50°, only 23 parts; but from this point the solubility increases up to 200 parts at a temperature of 212° F.

Substances dissolved in alcohol, ether, oils, and water are called simple solutions. The most important solvent is water; and when the substance is soluble in this liquid no other is used.

Elevation of temperature accelerates all chemical solutions: heat promotes the action of bodies on each other. The quantity of a dissolved body in a chemical solution is always the same. When bodies enter into chemical combination, they unite in definite and fixed proportions irrespective of temperature. Solutions are made either in test tubes, Florence flasks, evaporating dishes, or beakers, the size being regulated by the quantity of the substance to be dissolved.

Filters and Filtration.—Filtering-paper should be of uniform texture, and sufficiently porous to admit a free passage of the fluid. Filtration is purely a mechanical operation. The object is to separate the solid from the liquid, the former collecting on the filter, and the other passing through, which is called the filtrate. The filtering-paper is supported by a glass funnel, and the funnel may be supported by one of the rings of the retort stand. A filter is easily made by folding the paper in such a manner as to bring the four corners together, and, by opening one corner from the other three, a conical cavity is formed. The filter may now be placed in the funnel, taking care that it is the proper size. It should fit nicely against the sides of the funnel, and the substance to be filtered added by gently pouring it on the side. If projected violently into the centre of the filter, the paper is liable to be broken. Nothing short of actually doing the thing will give the proper knowledge. It would be just as reasonable to expect a man to learn the trade of a joiner by reading books on the construction of tools, as to expect a person to learn chemistry without proper apparatus; but this apparatus need not be expensive, if the student have (as he ought to have) some inventive and constructive power.

NON-METALLIC BODIES.

The metalloids or non-metallic bodies may be divided into four classes, each division being characterised by a great resemblance in their chemical properties: —

First Division.
Oxygen,
Sulphur,
Selenium.

Second Division.
Chlorine,
Bromine,
Iodine,
Fluorine.

Third Division.
Nitrogen,
Phosphorus.

Fourth Division.
Carbon,
Silicon,
Boron,
Hydrogen.

FIRST DIVISION.

OXYGEN.

Symbol, O; *Equivalent*, 8.

The non-metallic elements, or metalloids, constitute a large proportion of the material world. Oxygen gas was discovered by Priestley, in 1774, who called it dephlogisticated air. It was also discovered about the same time by Scheele, a Swedish chemist. The term oxygen was given to this gas by Lavoisier. This element is largely distributed throughout nature, both in the organic and inorganic kingdom. It constitutes eight parts out of every nine of water, which covers three-fourths of the globe. The solid crust of the globe is one-third oxygen, as will be seen when we consider the silica, lime, and alumina of which the mineral crust is composed.

The atmosphere is composed by weight of 23 parts of oxygen, and 77 parts of nitrogen. By volume it contains 79 parts of nitrogen, and 21 parts of oxygen.

Preparation.—Oxygen may be prepared in a variety of ways. By exposing to heat in a small glass retort the red oxide of mercury. In this substance the affinity which holds the oxygen to the mercury is so feeble that exposure to heat is sufficient to cause a separation into oxygen gas and metallic mercury. The mercury collects in small globules in the cold part of the retort, and the oxygen passes over as a colourless, inodorous, invisible gas: $HgO = Hg$ and O free. The chief interest connected with this method is that it was the one adopted by Priestley: it is seldom adopted by chemists, as being too expensive. It may also be prepared by exposing to a red heat, in an iron bottle or retort, the binoxide, or peroxide as it is sometimes called, of manganese, MnO_2. The oxide of manganese yields up half an equivalent of oxygen, reducing it to the sesquioxide, $Mn1\frac{1}{2}O$ which doubled makes Mn_2O_3. This is the method usually adopted when large quantities of oxygen are required, but the oxygen is seldom pure. Oxygen may also be prepared by distilling in a glass retort the binoxide of manganese and sulphuric acid. These materials should be well mixed, so that no dry place appears in that part of the retort where the materials are placed. If this is not attended to, fracture of the retort is likely to take place. Sulphuric acid will not unite with the binoxide of manganese, but it unites with the protoxide readily. The decomposition is easily illustrated: $MnO_2 + SO_3 = MnO,SO_3 + O$ free. Oxygen is also prepared by the decomposition of water by voltaic electricity: this is interesting as determining the composition of water, the hydrogen being liberated at the negative electrode, and the oxygen at the positive, in the proportion of double the volume of hydrogen to that of oxygen, which is the proportion in which these gases combine by volume to form water. The usual method of preparing oxygen is by heating a mixture of equal parts of the binoxide of manganese and the chlorate of potash, which should be well dried in a capsule. The manganese in this case undergoes no change, but the gas is given off at a much lower temperature than when the chlorate is used alone. In the preparation of gases the first part is a mixture of air, and should always be allowed to escape. The chlorate of potash is composed of 1 equivalent of potassium, 1 of chlorine, and 6 of oxygen; on the application of heat the whole of the 6 equivalents of oxygen is given off. Decomposition, $KO,ClO_5 = KCl + 6O$.

Oxygen, at all observed temperatures is a gas. It is not combustible; that is, it will not burn, but it is a powerful supporter of the burning of other bodies. It has no acid or alkaline reaction, although it enters into the composition of all oxygen acids and alkalies. Bodies which burn in air, burn with increased brilliancy in pure oxygen. A strip of wood, if only a spark remain, is instantly rekindled if plunged into a jar of this gas. Small iron wire, twisted into a spiral like a cork-screw, and the end dipped in sulphur, burns with beautiful little sparks. Phosphorus, with a white light of great brilliancy; sulphur, with a beautiful blue light; and charcoal burns with beautiful scintillations. In all these cases of combustion we have either acids or oxides. In the iron we have the black oxide, oxygen and iron, in the proportion of three equivalents of iron to four of oxygen, Fe_3O_4. In the phosphorus, phosphoric acid, PO_5 in white solid flakes, which dissolve in the water: in the sulphur, sulphurous acid, SO_2, which also dissolves in the water. The burning charcoal forms carbonic acid. In all these cases chemical combination takes place between the oxygen and the burning body ; and the weight of these bodies is increased in exact proportion to the oxygen consumed. All the common cases of combustion which take place in air are precisely the same as those which take place in oxygen. The action in pure oxygen is very energetic, owing to the absence of nitrogen. In the atmosphere these changes take place in a more subdued manner. The respiration of animals is, in effect, the same as common combustion. The substances in the blood burn away slowly by the aid of the oxygen introduced into the system. The activity of oxygen never ceases, although the phenomena of combination is not so marked as those just described. An ounce of iron rusting in the air, produces just the same amount of heat as if it had been burnt like the wire in oxygen. This slow combustion, unattended by light, extends over several years, while the same result could be produced by pure oxygen in a few seconds.

The decay of animal and vegetable substances is a process of slow oxidation. The offensive impurities of the air and water are destroyed in a great degree by a slow process of burning. The lungs of land animals, and the gills of fish, are adapted to absorb oxygen. The animal from the air, and the fish from the dissolved oxygen of the water. An animal confined in a given bulk of air dies as soon as the oxygen is consumed ; but if confined in the same bulk of pure oxygen, the circulation quickens, and fever and death are the results. Such is a brief outline of the relationship of oxygen to animal life; its action seems destructive, yet it is the sustainer of life, the mainspring of all vital action. Oxygen may be distinguished from every other gas, except nitrous oxide, by plunging into it a taper which has just been blown out; if the smallest part is incandescent, it immediately bursts into flame.

SULPHUR. BRIMSTONE.

Symbol, S; *Equivalent*, 16.

At ordinary temperatures this element is a pale yellow crystalline solid. It is often found in a free state. Most of the sulphur employed in this country comes from Sicily and other volcanic districts. In combination with iron, antimony, zinc, lead, and other metals, in union with oxygen, as sulphuric acid, it combines with lime, baryta, and magnesia, forming sulphates. It is present in small quantities in all animal and vegetable substances. It exists in the albumen of eggs. The blackening of silver spoons is owing to the formation of the sulphide of silver.

Sulphur melts at 230°, becoming a thin amber-coloured liquid; when heated to 430° it thickens like treacle. In this state it may be poured from a flask into cold water, when it retains for some time a soft, flexible condition, like gutta percha; this is the amorphous condition of sulphur. If it be raised to a temperature of 790° it boils; and if the vapour be conducted into a cold chamber it condenses in crystals known as sublimed sulphur or the flowers of sulphur.

Sulphur in many of its chemical relations bears a close analogy to oxygen. Many of the oxides of the metals have corresponding sulphides ; these sulphides often unite, forming compounds analogous in composition to salts. The roll sulphur of commerce is formed by running it when in a melted state into wooden moulds. There are probably seven oxygen compounds of sulphur —

Sulphurous acid	SO_2
Sulphuric acid	SO_3
Hyposulphurous acid	S_2O_2
Hyposulphuric acid	S_2O_5
Sulphuretted hyposulphuric acid	S_3O_5
Bisulphuretted hyposulphuric acid	S_4O_5
Trisulphuretted hyposulphuric acid	S_5O_5.

The two first only of these compounds will engage our attention.

When sulphur is burned in oxygen, sulphurous acid, SO_2, is formed. It may also be prepared by heating equal weights of sulphuric acid and mercury or copper turnings. The sulphurous acid comes off as a colourless gas. It must be collected by displacement as its great solubility in water will not admit of its collection at the pneumatic trough. The decomposition:—

$$2HO,SO_3 \text{ and } Hg \text{ or } Cu = HgO,SO_3 \text{ or } CuO,SO_3 + SO_2.$$

Sulphurous acid has bleaching properties, and is used for bleaching straw used in plaiting, and woollen goods. The plaiters in Bedfordshire suspend their plait in a tea chest, and set fire to some sulphur in an iron dish at the bottom ; the top of the chest is then covered over with a sack. If blue violets, or any coloured flowers, be suspended in a jar of this gas, the colour is entirely removed. Litmus paper, dipped in a solution of sulphurous acid, is first reddened, then bleached. It has also the power of arresting animal decomposition.

Sulphuric acid, SO_3.—This is a most important compound. It is extensively used in the manufacture of soda from common salt, also in the manufacture of chlorine, for bleaching, for dissolving indigo, dyeing, calico printing, and numerous other applications. The hydrated acid, HO,SO_3, has been known from the fourteenth century. It is now prepared by two methods:—from the distillation of the sulphate of iron, $FeO, SO_3 + 7HO$, green vitriol, and by the oxidation of sulphuric acid with the vapour of hyponitric acid, NO_4. The first method is that adopted on the Continent. The sulphate of iron is well dried, and placed in a stoneware retort, which is raised to a red heat. The sulphuric acid comes over in combination with water. This acid, when strongest, contains two equivalents of acid to one of water, $HO,2SO_3$. It is much stronger than the English acid. Sulphuric acid, when anhydrous, SO_3, is a solid, and may be obtained by gently heating the strongest acid in a retort provided with a receiver, surrounded by a freezing mixture. The vapour, on reaching the receiver, is condensed into white silky crystals. It has no acid properties; but, when thrown into water, it hisses like red-hot iron. On exposure to the air, it liquefies by the absorption of moisture. The method usually adopted in the manufacture of sulphuric acid depends on the formation of sulphurous acid, and its subsequent oxidation into sulphuric acid. Sulphur is kindled in a furnace freely supplied with air; the flame of the sulphur heats a crucible containing a mixture of nitre and sulphuric acid; these vapours are conducted into a long leaden chamber. The floor of the chamber is covered with water, and a jet of steam plays into it. At the end of the chamber an exit tube is provided for the spent gases. By this arrangement a constant supply of sulphurous acid, nitrous acid, water, and atmospheric air is provided. Two equivalents of sulphurous acid, $2SO_2$, take two equivalents of oxygen from one equivalent of hyponitric acid, NO_4, reducing the hyponitric acid to the binoxide of nitrogen, NO_2. This binoxide of nitrogen, in contact with the atmosphere, takes two equivalents of oxygen, returning to the state of hyponitric acid. This hyponitric acid, like the first quantity, again gives up two equivalents of oxygen to two equivalents of sulphurous acid. Water is in all cases essential to this change, which may be expressed thus:—

$$HO + NO_4 + 2SO_2 = 2SO_3 + NO_2 + HO.$$

This is a general outline of the theory of the process. A small quantity of hyponitric acid is sufficient to convert a large quantity of sulphurous acid into sulphuric. The hyponitric acid plays the part of carrier between the oxygen of the air and the sulphurous acid. At intervals the acid is drawn off from the floor of the leaden chamber, and is deprived of any excess of water by evaporating in shallow leaden pans. Sulphuric acid has a great affinity for water. Gases are frequently dried by being transmitted through this acid. The nitrate of baryta, or the chloride of barium, is a test for sulphuric acid, or any soluble sulphate. A white precipitate is formed, which is insoluble in nitric acid.

Sulphuretted hydrogen, HS, or hydrosulphuric acid. This compound of sulphur and hydrogen is very useful in the laboratory. It may be prepared by the action of dilute sulphuric acid on the protosulphide of iron. The apparatus for hydrogen may be used. The change may be expressed thus:—

$$\mathrm{FeS} + \mathrm{HO,SO_3} = \mathrm{HS} + \mathrm{FeO,SO_3}.$$

Sulphuretted hydrogen, at the common temperature, is a gas very soluble in water, having a most disagreeable odour, and very deleterious to animal life. If a jar of this gas is kindled, it burns a blue flame, depositing the sulphur. When mixed with chlorine, the whole of the sulphur is liberated. Cold water dissolves its own volume of this gas, and the solution is kept as a test. It soon spoils by absorption of the oxygen, and is best prepared as wanted. Jars of this gas may be collected over warm water; but the pneumatic trough, or any painted vessel, must not be used, as it will be blackened by the formation of a metallic sulphide. This gas is a frequent natural product. It occurs in the Harrogate waters, and is a frequent product of the decay of animal and vegetable matter. It has the properties of an acid; reddens litmus, and is called hydrosulphuric acid, which must never be confounded with hydrated sulphuric acid, $\mathrm{HO,SO_3}$. A piece of white blotting paper, dipped in a solution of acetate of lead, is immediately blackened by the smallest trace of this gas. This method is frequently adopted to test the purity of coal gas. When it occurs in apartments or drawers, it may be entirely destroyed with chlorine. Silversmiths usually burn their gas outside, owing to the sulphuretted hydrogen almost always present in small quantities in coal gas, blacking the silver by the union of the sulphur with that metal.

SELENIUM.

Symbol, Se; Combining Equivalent, 39·5.

This is a rare substance, analogous to sulphur, and forming seleniates, which are identical with the sulphates. It forms three compounds with oxygen:—

Oxide of selenium, SeO.—A gas, has the smell of horse radish, not acid, and slightly soluble in water.

Selenious acid, SeO_2.—This is analogous to sulphurous acid.

Selenic acid, SeO_3.—The hydrated acid resembles sulphuric acid.

* * *

SECOND DIVISION.

CHLORINE.

Symbol, Cl; *Combining Equivalent*, 35·5.

This element is very abundantly diffused throughout nature. In combination with the metal sodium, it forms common salt. It is never found in its free or uncombined state. At ordinary temperatures, it is a gas of a greenish colour; and in very small quantities has an odour like sea water; but in large quantities, it produces violent coughing and irritation. It is prepared by mixing, to the consistency of cream, strong hydrochloric acid, and the black oxide of manganese. Care should be taken that every part is well mixed, so that no dry place appears inside the retort or flask, or fracture is most likely to result. A gentle heat should now be applied. When the gas comes over in abundance, it should be collected over warm water, as cold water absorbs about twice its volume; or it may be collected by displacement, like carbonic acid. The reaction is easily understood:—

$$2HCl + MnO_2 = 2HO + MnCl + Cl.$$

Chlorine has little affinity for oxygen; it combines with hydrogen and the metals. If a lighted taper be immersed in a jar, it burns with a dull red light, with an abundant evolution of carbon, in the form of smoke. If a piece of blotting-paper be soaked in turpentine, and placed in a jar of chlorine, it bursts into flame, combining with the hydrogen of the turpentine, and liberating the carbon. In both these cases, the chlorine combines with the hydrogen to form hydrochloric acid, to the exclusion of the carbon, which falls as soot—a very good illustration of the superior affinities of chlorine for hydrogen.

Phosphorus combines with chlorine spontaneously, but the flame is very feeble. Copper-leaf, powdered antimony, and arsenic, all undergo combustion in chlorine. If equal volumes of hydrogen and chlorine be introduced into a suitable apparatus, and an electric spark sent through the mixture, combination takes place with great violence. The direct rays of the sun produce a similar union. The most remarkable property of this gas is its power of destroying organic colouring matters. A solution of indigo, which is not affected by strong sulphuric acid, is changed into a brown liquid by chlorine, to which the original colour cannot be restored. In a perfectly dry state, chlorine has no bleaching properties. When in solution, it may be poured into test-glasses, containing infusions of litmus, purple cabbage, turmeric, carmine, cochineal, &c. The colour of these substances will be removed. When bleaching is practised on a large scale, it is generally used in combination with lime, called bleaching powder. The outline of this process can be easily imitated. The bleaching powder is mixed with warm water, and the solution filtered. Any piece of coloured calico may be taken, providing its tint is due to some organic colouring matter. It may then be washed in the filtered solution. The liquid bleaches but slightly; but if it be removed from the bleaching bath, and rinsed in a bath of dilute sulphuric acid, and this process carried on a few times alternately, the whole of the colour is removed. The chlorine, when in combination with the lime, is comparatively inactive; but when dipped in dilute acid the lime combines with the acid, and the chlorine in the presence of water is set free, and manifests its bleaching properties. An experiment illustrating this property may be performed in the following manner. A piece of cotton, dyed with madder, or any organic colouring matter, may have any pattern or device stamped upon it, or painted with a mixture of gum and tartaric acid. The cloth is dried, and afterwards rinsed, in a hot solution of bleaching powder. A white pattern on a coloured ground is developed in consequence of the acid mixed with the gum liberating the chlorine, and destroying the colour in those parts upon which the pattern was printed. The other parts of the cloth are not sensibly affected.

Chlorine is used as a disinfectant; it destroys unpleasant and offensive odours. The chlorine is usually evolved from chloride of lime, which is dissolved in hot water, and a little vinegar or dilute sulphuric acid added to liberate the gas.

Hydrochloric Acid, HCl.—This is a most important acid. Hydrochloric acid gas is prepared by heating in a properly-arranged flask a mixture of common salt, a little water, and sulphuric acid. It must be collected by displacement, or over mercury. It is colourless, fumes when in contact with the atmosphere, and is very soluble in water; one cubic foot of that liquid dissolving about 418 cubic feet of gas. A solution of this gas is the acid usually known as muriatic or hydrochloric acid. It may be prepared by mixing equal weights of dry common salt and sulphuric acid in a flask, and applying heat; the gas after being purified by passing it through a wash-bottle, and from thence into distilled water, by which the gas is absorbed. The reaction is easily understood:—

$$NaCl + HO,SO_3 = NaO,SO_3 + HCl.$$

A mixture of nitric and hydrochloric acid is called *aqua regia*, from its property of dissolving gold.

The presence of any soluble chloride, or hydrochloric acid, may be detected by adding a solution of the nitrate of silver, when a white curdy precipitate is produced, insoluble in nitric acid:—

$$NaCl + AgO,NO_5 = NaO,NO_5, \text{ soluble}: AgCl, \text{ insoluble.}$$

There are five compounds of oxygen with chlorine:—

Hypochlorous acid	ClO.
Chlorous acid	ClO_3.
Hypochloric acid	ClO_4.
Chloric acid	ClO_5.
Perchloric acid	ClO_7.

ClO_3, ClO_4, and ClO_7 are obtained from chloric acid. The other compounds are obtained by the action of chlorine on metallic oxides.

Hypochlorous acid, ClO, is a yellow gas, has very powerful bleaching properties, and may be prepared by adding the powdered red oxide of mercury to a bottle of chlorine containing a little water. The bottle should be briskly shaken for a few moments. The oxygen of the mercury combines with the chlorine to form ClO, which dissolves in the water, and forms a compound called the oxychloride of mercury; thus—

$$2HgO + 2Cl = HgCl,HgO + ClO.$$

Chlorous acid, ClO_3.—This compound is very difficult of preparation, and not very important.

Hypochloric acid, ClO_4, sometimes called the peroxide of chlorine.—This acid may be prepared by carefully mixing concentrated sulphuric acid and the chlorate of potash into a thick paste. After it has cooled it should be cautiously introduced into a green glass retort, and very slowly heated by placing the retort in water; and, gently raising the temperature of the water by a lamp or other source of heat, a green-yellow gas comes off, which is hypochloric acid, ClO_4. It must be collected by displacement, as it is very soluble in water, and mercury decomposes it.

This compound is very explosive. At a temperature below boiling water it is resolved with great violence into oxygen and chlorine; hence the care required in its preparation. Decomposition:—

$$3(KO,ClO_5)+3(HO,SO_3)=2ClO_4+ClO_7+ \\ 3(KO,SO_3)+3HO.$$

If a mixture of chlorate of potash and loaf sugar be placed on a plate and touched with a glass rod previously dipped in sulphuric acid, it is instantly set on fire. In this experiment the hypochloric acid,

ClO_4, is disengaged with such violence as to cause combustion.

If half a teaspoonful of chlorate of potash and a few chips of phosphorus be placed in a deep test or ale glass nearly filled with water, and a little sulphuric acid by a small funnel which reaches to the bottom of the glass, the phosphorus will burn under water from the oxygen of the disengaged hypochloric acid.

Chloric acid, ClO_5, forms in combination with bases salts of considerable importance, especially the chlorate of potash. If a current of chlorine be sent through a moderately warm solution of caustic potash, and the liquid, when saturated with the gas, concentrated by evaporation, crystals of chlorate of potash are obtained. In this reaction a part of the potash, KO, is decomposed to form chloric acid and the chloride of potash:—

$$6KO + 6Cl = KO,ClO_5 + 5KCl.$$

From the chlorate of potash chloric acid is obtained, by boiling with a solution of hydrofluosilicic acid. The process is tedious, and requires care.

Perchloric acid, ClO_7, is obtained from the perchlorate of potash, KO, ClO_7.

The chlorates in solution give no precipitate with the nitrate of silver. When heated they evolve oxygen; and if treated with sulphuric acid they evolve the characteristic yellow gas as mentioned under hypochloric acid.

BROMINE.

Symbol, Br; *Equivalent*, 78·26.

This element is found in sea-water at ordinary temperatures. It is a brownish red liquid, having a disagreeable odour, closely resembling chlorine and iodine in character. It is found in combination with magnesium in sea water, as the bromide of magnesium, $Mg'Br$. It is obtained by agitating ether with any solution of bromine. After the less soluble salts have been separated, a stream of chlorine is then sent through the solution, which decomposes the bromide of magnesium, the ether dissolving the liberated bromine. The liquid is transferred to a retort with a receiver, and heat applied, when the bromine condenses in the cold part of the retort and receiver. Bromine, when associated with water has bleaching properties.

Hydrobromic acid, HBr, is analogous to hydriodic acid, and has nearly the same properties, and may be prepared by substituting bromine for iodine, in the preparation of hydriodic acid.

There is only one known oxide of bromine, bromic acid, BrO_5, which is prepared by adding bromine to a solution of caustic potash, until all alkaline properties are destroyed. The change is the same as that described under chloric acid.

IODINE.

Symbol, I; *Combining Equivalent*, 127.

At ordinary temperatures, iodine is a bluish-black, metallic looking substance, having the appearance of scales of black lead. It occurs chiefly in combination with sodium and potassium in sea water. It is a constituent of marine plants and animals. It is chiefly prepared from the ashes of sea weeds, called kelp. The ashes are mixed with water, filtered and evaporated, until the solution is reduced to a very small quantity: the salts of sodium, potassium, and magnesia being removed as they crystallise. The mother liquor left after the removal of these salts, contains the iodine in combination with magnesia and iodine. This is placed in a leaden retort with sulphuric acid and the binoxide of manganese. On the application of heat the iodine distils over as a dark purple vapour, which is condensed. This process may be imitated by placing in a retort connected with a receiver, a solution of the iodide of potassium, a little powdered binoxide of manganese, and sulphuric acid: on the application of heat, a beautiful violet vapour passes over, which condenses in the cold part of the retort and receiver in small shining scales. Decomposition:—

$$KI + MnO_2 + 2HO,SO_3 = HO,SO_3 + MnO,SO_3 + HO + I.$$

Iodine stains the skin; it acts violently on the animal system, and is used in medicine.

When iodine is free, it produces a splendid blue colour in contact with starch. If any soluble iodide be first liberated with a little chlorine water, it is then capable of producing the characteristic colour with starch.

Iodine forms with oxygen two acids, iodic acid, IO_5, and periodic acid, IO_7.

In combination with hydrogen it forms an acid known as hydriodic acid, HI. This acid is a gas, but it may be obtained in solution by passing sulphuretted hydrogen through water in which powdered iodine is suspended. In this case, the sulphur of the sulphuretted hydrogen is liberated, and the hydrogen combines with the iodine, forming hydriodic acid, $HS + I = HI + S$. The liquid is warmed, filtered, and a colourless solution of the acid is obtained, which soon becomes of a brown colour on exposure to air, by its decomposition into free iodine and water, $HI + O = HO + I$. Iodine dissolves freely in alcohol. If a polished silver plate be held over the vapour of iodine, it first becomes yellow, then violet, and then a deep blue, owing to a combination of the iodine with the silver. The iodide of silver thus formed is decomposed by light.

FLUORINE.

Symbol, F; *Equivalent*, 18·70.

Owing to the powerful affinities of this element for the metals, also for silicon and hydrogen, it has never been isolated. Its most frequent combination is with calcium, as the fluoride of calcium, CaF, or fluor-spar. It exists, also, in a state of combination, in the bones and teeth of animals. So far as it has been examined, it is said to be a gas, like chlorine.

There is no known oxygen compound of fluorine, but in combination with hydrogen, it forms an acid known as hydrofluoric acid, HF. To obtain this compound, the powdered fluoride of calcium is heated with concentrated sulphuric acid in a platinum or lead retort connected with a receiver of the same metal. Kept cool by being surrounded with a freezing mixture, a volatile, colourless liquid, is obtained, which emits suffocating white fumes in the air. It unites with water, and very dilute solutions attack glass. It has, therefore, to be kept in leaden, or gutta-percha bottles.

The decomposition is easily explained:

$$CaF, HO,SO_3 = HF + CaO,SO_3.$$

If the concentrated acid be dropped upon the skin, it occasions ulcers of a very malignant and dangerous kind. Hydrofluoric acid is used to etch glass, or engrave letters. The glass to be operated upon is first covered with bees' wax, and the lines to be etched, scratched through with a needle. The glass is then exposed to the action of the vapour of hydrofluoric acid, or the acid is poured upon it in a liquid state. Where the glass is unprotected with wax, it is corroded. It is in this manner that thermometers, and other glass vessels, are graduated. The power of the acid to etch glass may be observed by having a leaden saucer, into which some powdered fluor-spar and sulphuric acid is placed. The glass to be etched is first covered with a coating of wax, and the device drawn upon it with a pen-knife so as to remove the wax from those parts to be etched. The glass which should cover the leaden saucer is now placed on the top, with the waxed side downwards. The saucer should now be gently heated, care being taken that the temperature does not rise sufficiently to melt the wax. In a few minutes the glass may be removed and cleaned with a little hot turpentine, when it will be found to be etched.

THIRD DIVISION.

NITROGEN.

Symbol, N; *Combining Equivalent*, 14.

This gas is without taste, odour, or colour, slightly lighter than common air, and distinguished from all the other gases by its negative properties. It forms by volume $\frac{4}{5}$ of the atmosphere, and enters into a variety of combinations. Nitrogen exists in the vegetable kingdom. It is probably derived from the ammonia of the atmosphere. It is a constituent of muscular tissue to the extent of about 17 per cent. It is not found in the mineral crust of the earth, except in some kinds of coal.

Nitrogen may be prepared by burning a small piece of phosphorus in a confined portion of common air under a gas jar standing over water, by which phosphoric acid is produced, and the nitrogen liberated. The white solid flakes of phosphoric acid are dissolved in the water over which the jar is placed. It must stand till the jar has become clear. Theory of change:—

$$\text{Air}\left\{\begin{array}{l}O\\N\end{array}\right. \quad \text{Phosphoric acid, } PO_5.$$
$$P$$

By passing chlorine through a solution of ammonia:

$$3Cl + NH_3 = 3HCl + N \text{ free.}$$

By mixing iron filings and sulphur with water and allowing the mixture to remain under a gas jar for twenty-four hours. The oxygen of the atmosphere is absorbed and the nitrogen set free. By treating nitric acid with lean butcher's meat in a retort. This change is not well understood.

The atmosphere, which is the great storehouse of nitrogen, is composed of a mixture of nitrogen and oxygen, with variable proportions of carbonic acid, aqueous vapour, and traces of ammonia.

The analysis of common air may be effected in the following manner:—Introduce, by means of a wire, into a graduated tube standing over water, a piece of phosphorus. The known quantity of air can be determined by observing the water point in the graduated tube. The arrangement should now be left for twenty-four hours. The whole of the oxygen will be absorbed by slow combination with the phosphorus. The remaining gas may now be read off, which is nitrogen. Another plan is to mix air with hydrogen and explode by an electric spark, the hydrogen and oxygen combine, and the remaining volume of gas read off. Nitrogen is not combustible, nor is it a supporter of combustion or respiration, it is destitute of any acid or alkaline reaction. It is estimated that six tons of air pass through a blast furnace every hour. While the oxygen is most active in forming new compounds the nitrogen escapes unchanged. Its chief function is to modify the violent action of the oxygen, and render the atmosphere suitable to the wants of life.

There are five compounds of oxygen and nitrogen, which may be taken as a good illustration of the law of multiple proportions in chemistry :—

Protoxide of nitrogen . .	NO
Binoxide of nitrogen . .	NO_2
Nitrous acid	NO_3
Hyponitric acid . . .	NO_4
Nitric acid	NO_5

The protoxide of nitrogen is sometimes called laughing gas, from the peculiar exhilarating effects produced when inspired. It is prepared by heating moderately in a retort the nitrate of ammonia. Four ounces of salt produce one cubic foot of gas. It should be collected over warm water, and care should be taken that the salt is pure. The change which takes place is illustrated by the following equation :—

$$NH_3, NO_5 = 3HO + 2NO.$$

Nitrous oxide is a supporter of combustion. The inhalation of this gas is sometimes attended with serious results.

The binoxide of nitrogen may be prepared by introducing into the hydrogen apparatus some copper turnings, a little water, and nitric acid, added by the funnel. The gas may be collected over water. The gas obtained in this manner is colourless and transparent; but in contact with air or oxygen it produces deep red fumes, which are sufficient to distinguish it from all other gases. This may be beautifully illustrated by passing some binoxide of nitrogen into a gas jar containing oxygen, standing over the pneumatic trough, or allowing the fumes to pass into the atmosphere:—

$$3Cu + 4NO_5 = 3(CuO,NO_5) + NO_2.$$

Nitrous acid, NO_3.—When four volumes of dry binoxide of nitrogen are mixed with one volume of dry oxygen, and exposed to as low a temperature as can be obtained by freezing mixtures, a thin mobile liquid of a green colour is produced with a vapour of orange-red. This is nitrous acid, NO_3.

Hyponitric acid, NO_4.—If the nitrate of lead be heated in a retort of hard glass it is decomposed into the protoxide of lead, oxygen, and hyponitric acid. It may be collected in a receiver surrounded with a freezing mixture. Great obscurity hangs over these two compounds of nitrogen and oxygen. The red fumes produced when the binoxide of nitrogen comes in contact with oxygen are chiefly due to the formation of hyponitric acid. Other oxygen compounds are also formed.

Nitric acid, NO_5.—Known as aquafortis, and one of the most important compounds of nitrogen. In hot dry climates, the soil is sometimes covered with a white, downy looking salt. When this salt is collected, dissolved, filtered, and crystallised, it produces the nitre or saltpetre of commerce: it is a compound of potash and nitric acid. To obtain nitric acid from this salt, mix equal weights of sulphuric acid and nitrate of potash in a glass retort, cover the receiver, which may be a Florence flask, with a wet cloth, and apply heat. Red fumes at first rise in the retort, and pass to the receiver; these fumes gradually disappear, and a pale yellow liquid collects in the receiver, which is nitric acid. Decomposition:—

$$KO,NO_5 + HO,SO_3 = KO,SO_3 + HO,NO_5.$$

On a large scale earthen retorts and condensing vessels are employed, and the nitrate of soda substituted for the nitrate of potash.

Pure nitric acid is a colourless liquid. It fumes when exposed to air. It is partially decomposed by light. It is intensely acid, and reddens vegetable blues. It stains the skin and nails yellow, and is used for producing yellow colours on woollen fabrics. It is also used for assaying, and as a solvent for tin in calico printing. It is a powerful oxidising agent, attacking most of the metals with great violence.

Nitric acid forms a series of salts, all of which are soluble in water. After a thunderstorm, nitric acid, in small quantities, has been detected in rain water. It is supposed by some, to be produced in the air by electricity.

When nitrates are heated with an excess of hydrochloric acid they are converted into chlorides, with the formation of nitric acid and the disengagement of free chlorine,

$$KONO_5 + 2HCl = KCl + 2HO + NO_5 + Cl.$$

When paper is soaked in a solution of potash, and dipped in nitric acid it becomes touch-paper. Cotton, soaked in nitric acid, washed and dried at a temperature not exceeding 120°, forms gun-cotton. The low temperature at which it explodes has prevented its substitution for gunpowder.

PHOSPHORUS.

Symbol, P; *Combining Equivalent,* 32.

This substance was discovered by the alchemists two hundred and fifty years ago. Phosphorus, at ordinary temperatures, is a solid, waxy-looking substance, easily cut with a knife. It is poisonous. In the dark it shines with a faint blue light, and is highly combustible, often taking fire with the slightest touch or friction. Phosphorus is always kept under water, and great care is required in experiments with it.

Phosphorus is prepared from bones; they are first well calcined, reduced to powder, and mixed with two thirds their weight of dilute sulphuric acid; after some time the mixture is filtered. The sulphate of lime which has been formed by the sulphuric acid is a very insoluble compound and easily separated on the filter. The liquor is then evaporated to the consistence of treacle, mixed with charcoal powder, and transferred to an iron vessel which is raised to a high temperature till all moisture is expelled. After this mixture has been thoroughly dried it is placed in a stoneware retort, from which a large tube dips into water contained in the receiver. The phosphorus distils over in a soft condition, and is moulded in hot water into cylindrical sticks. The change which takes place in the retort is easily understood. Phosphoric acid contains one equivalent of phosphorus and five equivalents of oxygen. The oxygen combines with five equivalents of carbon, forming carbonic oxide, which escapes:—

$$PO_5 + 5C = P + 5CO.$$

The phosphates play a very important part in the solid structure of animals, giving stiffness and rigidity to the bones. The dried skeleton of a man weighs about 10 or 12 lbs., and contains about 2 lbs. of phosphorus. Traces of phosphorus have been found in the brains, flesh, and blood of all animals. There is an important modification of phosphorus discovered by Professor Schrotter, called allotropic phosphorus, in which state the active properties of ordinary phosphorus are suspended; it is of a brick-red colour, insoluble in the bisulphide of carbon, does not become luminous in the dark, has no tendency to combine with oxygen, is not poisonous, and may be carried about without fear.

Phosphorus is largely used for the manufacture of lucifer matches. As the phosphorus would be likely to ignite by exposure to the atmosphere, it is mixed up with glue or gum water. The ends of the matches are first dipped in sulphur, then into the phosphorus solution, and afterwards into a thick mixture of chlorate of potash or nitre, or some substance rich in oxygen; they are then removed to a stove and carefully dried. About 200,000 to 250,000 lbs. of phosphorus are used annually in London in the manufacture of these matches. When the surface of the match is broken by friction the phosphorus first takes fire; this inflames the sulphur, which ignites the wood: the cracking noise is owing to the chlorate of potash.

There are three well known compounds of phosphorus —

Hypophosphorous acid . . .	PO
Phosphorous acid . . .	PO_3
Phosphoric acid . . .	PO_5

The first of these compounds is not of much interest; it is a powerful deoxidising agent, and is rather difficult of preparation.

Phosphorous acid is formed by the slow combustion of phosphorus in a limited supply of air.

Take a glass tube, about one foot long and half an inch in diameter; at about two inches from one end bend the tube at an obtuse angle, drawing the end of the shorter limb out into a conical shape, leaving an aperture about the size of a pin's point. Into the bent end of this tube, at the angle, place a small piece of phosphorus carefully dried by folding it in blotting-paper. Gently heat the tube over where the phosphorus lies by means of a spirit lamp, when the phosphorus will take fire and burn with a greenish flame. Only a small quantity of air enters the tube through the small aperture. The acid condenses as a white powder in the upper cold part of the tube. This is solid anhydrous phosphorous acid. It may also be prepared in combination with water from the teriodide of phosphorus, PI_3. To prepare this iodide of phosphorus, place a small piece of dry phosphorus in a test-tube, cover it with about eleven times its weight of iodine; the two bodies combine energetically without any application of heat. This is a singular experiment, in which two solid bodies combine chemically. After the action is over, water is added, when the following decomposition takes place:—$PI_3 + 3HO = PO_3 + 3HI$; phosphorous acid and hydriodic acid. The liquid is heated till all the hydriodic acid has volatilised, and a solution of phosphoric acid is left.

Phosphoric acid, PO_5, is obtained when phosphorus is burned under a bell jar or in oxygen. When anhydrous, it is like snow, but, in consequence of its attraction for water, it soon dissolves, and, when it once becomes the hydrated acid, the water cannot be separated. Phosphoric acid may be prepared by adding little bits of phosphorus to nitric acid in a retort provided with a receiver: by this process the glacial phosphoric acid of the pharmacopœia is prepared; it is rather a troublesome and dangerous method.

Phosphoric acid may be obtained by adding sulphuric acid to bone earth, precipitating with the carbonate of ammonia, and separating the lime salt by filtration, then igniting and evaporating in a platinum crucible the sulphate and phosphate of ammonia; hydrated phosphoric acid remains behind. The most advantageous manner is the first process; namely, to burn phosphorus in a stream of dry air.

Phosphoric acid is not poisonous; it is, however, a powerful acid, reddens vegetable blues, and unites with bases to form a large and important class of salts.

Phosphuretted Hydrogen, PH_3.—This gas has a most disagreeable odour, resembling garlic; it is poisonous, slightly soluble in water; it precipitates several of the metals, but has no acid or alkaline properties. This gas may be prepared by heating in a small retort hydrated phosphorous acid, which is decomposed into phosphuretted hydrogen and hydrated phosphoric acid —

$$4(3HO,PO_3) = PH_3 + 3(3HO,PO_5).$$

It is more commonly prepared by boiling in a small retort sticks of caustic potash, water, and phosphorus; the retort should be nearly filled, and provided with a long tube which should dip into water from the commencement of the operation, and heat gently applied. The gas soon comes off with a slight explosion in the retort, attended with the rushing of the water along the tube. It is to prevent any accident that the tube should be long enough to prevent the water entering the retort: as soon as the oxygen is consumed all danger is over. As the gas escapes through the water into the air, it takes fire and changes into wreaths of white smoke. The preparation of this gas is a very beautiful experiment.

The presence of the potash determines the decomposition, similar to dilute sulphuric acid when in contact with zinc or iron.

The change may be represented thus —

$$4P + 3KO,HO = PH_3 + 3(KO,PO).$$

* * *

FOURTH DIVISION.

CARBON.

Symbol, C ; *Equivalent*, 6.

This element, in one state or another, has been known from the earliest time. It occurs in nature as the diamond, graphite, or black lead. Plumbago, charcoal, coke, bone black, soot, and lamp black, are artificial varieties of carbon. In combination with oxygen it forms two compounds, carbonic oxide and carbonic acid. The latter compound unites with potash, lime, soda, magnesia, forming carbonates, which constitute a large portion of the mineral crust of the globe. The compounds of carbon with hydrogen are all highly combustible. In combination with nitrogen and oxygen it enters largely into the composition of all animal and vegetable substances. The purest form of carbon is obtained by exposing animal or vegetable substances to heat, carefully excluding the atmosphere during the operation. This process, in the case of wood, is usually performed in large mounds covered with earth, so as to limit the supply of air, and allow the combustion to proceed slowly. With animal substances it is effected in retorts. The specific gravity of the charcoal varies with the wood from which it is manufactured.

Charcoal is very combustible, and is extensively used on the Continent for fuel. When perfectly pure it burns without flame. It is a very indestructible and unchangeable substance. Wooden piles driven into the earth are usually charred to prevent decay. Charcoal has the power of condensing in its pores different gases. It has been found to absorb ninety times its bulk of ammonia, and thirty-five times its bulk of carbonic acid. It is this property which makes it useful as a deodoriser. Charcoal when newly prepared restores meat slightly tainted, either by being boiled with it or rubbed over it. It removes offensive odours, and checks putrefaction. Rats buried in powdered charcoal yielded no unpleasant effluvia when opened after three months. Water is purified by passing it through filters constructed of layers of charcoal and sand. Black crayons are made from the charcoal of the willow. The charcoal used in the manufacture of gunpowder is made from dogwood or elder. Common bituminous coal is not carbon: it is a combination of many complex unions of carbon and hydrogen. Heat expels the volatile parts, and we have left coke, which is a mixture of carbon, with small quantities of metallic oxides. In the manufacture of gas these volatile products are purified and used for purposes of illumination.

Plumbago or black lead was formerly thought to be a carbonate of iron. The best specimens are altogether free from iron. Lead is never present in black lead. Plumbago is used for the manufacture of pencils. That of Borrowdale in Cumberland was the best, but the vein is now exhausted. Most of the ordinary pencils are manufactured from a paste made of powdered plumbago, antimony, and sulphur. Ivory and bone black are varieties of charcoal which result from the heating of bones in a retort, till all the volatile products are evolved. This bone black, however, only contains about 10 per cent. of carbon. The remainder is chiefly phosphate and carbonate of lime. Ivory black is used as ink in printing from copper and steel engravings. The preparation of bone black is now an important branch of chemical manufacture, as it is largely used in sugar refining, owing to its bleaching or decolorising properties.

It may not be altogether out of place to say a few words on sugar refining. If the saccharine juices of any vegetable substance be allowed to remain at rest, especially if the temperature be high, fermentation commences, and the sugar is changed into alcohol; it is therefore desirable to convert it into a crystalline state as speedily as possible. The juice of the sugar cane is raised to a temperature of about 130°, and lime water added. The whole mass is now raised in temperature, but not boiled. A thick scum rises to the surface. When this scum breaks it is either taken off, or the juice is drawn off by a tap at the bottom of the copper. By this continual heating the water evaporates, and the liquid thickens. Sugar cannot be heated in the presence of acids or alkalies without undergoing decomposition, and molasses may be considered as the product of this decomposition. After a sweet solution of sugar and water has been obtained, a portion of lime water is added, and bullock's blood. It is said that the lime is employed to neutralise acidity, but this is a mistake: no acid exists in unfermented sugar. The proper use of lime water is this: the colouring matter of some varieties of sugar unites with lime, and forms a precipitate which gets enveloped in the albumen of the blood, and rises to the surface in scum, which is removed. The solution is now allowed to escape into a filtering apparatus. This apparatus consists of a number of bags thrust one into the other, and fastened to the bottom of a tray, into which the sugar flows. As the liquor flows through the filter bags, a portion is examined from time to time, as to its brightness. When the solution has become clear it is of a dark port wine colour. The solution is now passed through bone black or animal charcoal. This decolouring property of bone black may be shown by passing a litmus solution through a charcoal filter. The water passes through quite clear. The whole theory of this decolorising effect is however involved in some obscurity. In sugar refining, a large cylinder of iron, six or eight feet in diameter, and twelve or fifteen feet high, is packed with granulated charcoal, within two or three feet of the top. The solution percolates through, and is drawn off at the bottom. The sugar solution which entered was of a dark colour; it passes from the filter as colourless as water. In time, however, the solution becomes tinted, the operation is stopped, and the charcoal washed. This is a brief outline of the process of sugar refining. Animal charcoal, has also deodorising properties.

Lamp black results from the imperfect combustion of tar, oil, resin. The greater part of the carbon is carried up in a state of finely-divided powder. When these particles come in contact with a cold body, they form a deposit of carbon, which may be scraped off. This forms the lamp-black of commerce. The operation is performed with a limited supply of air in a high chimney, containing a number of plates, upon which the lamp-black collects like soot in our own chimneys. Diamonds are carbon in a crystallised state. They were formerly obtained in India and the island of Borneo, but for the last hundred years they have been brought from Brazil. If a diamond be burned in oxygen, the sole result is carbonic acid, the same as if we had burnt a piece of charcoal. It is the opinion of some, that the diamond is of vegetable origin. Within the crystalline structure of many diamonds, the remains of insects are seen. Sir D. Brewster thinks the diamond a drop of fossilised gum.

For diamonds of moderate size, and the same quality, the value is estimated in this manner:—The weight is given in carats. These carats are subdivided into halves, quarters, eighths, sixteenths, and thirty-second parts. If four diamonds weigh 1, 2, 3, and 4 carats each, their respective values would be as 1, 4, 9, and 16: their value is in the ratio of the squares of their respective weights. The koh-i-noor weighed 186 carats: since it has been recut, it weighs $102\frac{3}{4}\frac{1}{16}$ carats.

There are two well known compounds of oxygen and carbon, carbonic acid, and carbonic oxide. Carbonic acid is present, in small quantities, in the atmosphere; it occurs in combination with lime, and magnesia; it is a constituent of all marbles, shells, and coral. Carbonic acid is present in all spring water; it is evolved from volcanoes, and is a constant product of respiration. Fermenting liquids, decaying animal and vegetable substances, and the combustion of coal, wood, and gas, are continual sources of carbonic acid at the ordinary temperature and pressure. It is a gas having slightly acid properties; under pressure it may be liquefied and solidified. Carbonic acid may be prepared synthetically by burning carbon in oxygen, or by decomposing a carbonate. Carbonate of lime is generally used for this purpose. When carbon is burnt in oxygen no change of volume takes place, although there is an increase of weight: 8 of oxygen combining with two equivalents of carbon, $8 + 6 + 6 = 20$. Lime water is a test for carbonic acid,—a milky, insoluble precipitate of carbonate of lime is formed when carbonic acid is passed through a solution of lime water.

Any carbonate yields carbonic acid when treated with any of the stronger acids. We use for this purpose hydrochloric acid and pieces of marble, which is a carbonate of lime, and a little water. The decomposition is as follows:—

$$CaO,CO_2 + HCl + HO = 2HO + CaCl + CO_2.$$

We may illustrate it in another way. Let us suppose the base to be an oxide of calcium, CaO (lime). The oxygen of the base unites with the hydrogen of the acid, and forms water. In all cases when hydrogen acids unite with an oxygen base, water is formed.

Carbonic acid is heavier than common air; it may be poured from one vessel to another; it is not a supporter of combustion or respiration, nor is it combustible. It does not extinguish flame, for the same reason as nitrogen, by excluding oxygen. For if we mix one volume of oxygen with four of carbonic acid, it does not support combustion, although the quantity of oxygen is as great as in common air. If a living animal be immersed in an atmosphere of nitrogen, it dies the same as if it were immersed in water. The nitrogen kills by suffocation; carbonic acid kills by poison: all the other conditions of life may be present, but the inhalation of carbonic acid produces death. In winter, people burn charcoal in braziers,—a most unhealthy and dangerous practice where there is not a proper exit for the carbonic acid produced. During the fermentation of beer or wine, carbonic acid is evolved. During respiration, it is given off. To illustrate this, take a glass tube and blow into lime-water: the water becomes milky, owing to the formation of carbonate of lime.

Carbonic Oxide.—The blue lambent flame sometimes seen on the top of brick-kilns and coke-ovens, and in our own fireplaces, is carbonic oxide (CO). This compound contains one equivalent less of oxygen than carbonic acid. It may be prepared by passing carbonic acid through a red-hot iron tube. Thus,—$Fe + CO_2 = FeO + CO$. It may also be prepared by passing carbonic acid through a porcelain tube filled with charcoal. The carbonic acid, CO_2, uniting with another equivalent of carbon, becomes converted into two equivalents of carbonic oxide, $CO + CO$. A similar change to this takes place in charcoal and coke fires: the charcoal combines with the oxygen, forming carbonic acid, CO_2; it passes through the red-hot part of the fire, and is converted into carbonic oxide, CO, and burns at the surface with a blue flame, which is sometimes said to be a sign of frost. A low temperature favours the production of this compound. Carbonic oxide is generally prepared by heating sulphuric acid with oxalic acid. As soon as the temperature is sufficiently raised, the gas comes off rapidly. The decomposition which takes place may be represented thus:—oxalic acid contains two equivalents of carbon, three of oxygen, and one of water, C_2O_3HO: this equals one equivalent of CO_2, one equivalent of CO, and one of HO. When heated, the sulphuric acid unites with the equivalent of water, and the carbonic acid, CO_2, and carbonic oxide, CO, come off together. By passing the gases through a solution of caustic potash, the carbonic acid is absorbed, and the carbonic oxide, CO, passes over, which may be collected at the pneumatic trough in the usual way.

This gas is invisible, and lighter than common air. It is highly combustible, but does not support either combustion or respiration. Dumas has shown that this gas is a hundred times more poisonous than carbonic acid,—a fact which till very lately was unknown.

The compounds of carbon and hydrogen, are very complex; they belong properly to organic chemistry. We shall describe two of these compounds.

Heavy carburetted hydrogen, C_4H_4, generally called olefiant gas, because it forms an oily compound when mixed with chlorine. It is lighter than common air. This gas is prepared by mixing alcohol with five or six times its weight of sulphuric acid, HO,SO_3. The acid should be added a little at a time, as it is easier to measure liquids than to weigh them; one of alcohol and three of sulphuric acid, HO,SO_3 is the proportion. The flask should be shaken so as to thoroughly mix the liquids.

$$\text{Alcohol, } C_4H_6O_2 = C_4H_4 + 2HO.$$
$$C_4H_6O_2 + HO,SO_3 = C_4H_4 + HO,SO_3 + 2HO.$$

Properties.—Olefiant gas is highly combustible, but not a supporter of either combustion or respiration: when burnt it forms carbonic acid, CO_2, and water, HO. Its combustibility may be shown by burning some from a jet in a jar half full of chlorine, Cl. If a light be applied to a mixture of these gases, HCl is formed, and the carbon liberated in a finely divided state.

Light carburetted hydrogen, CH_2.—This gas is disengaged abundantly from coal mines and from stagnant water. The miners call it fire damp; it is also called marsh gas. When mixed with common air it forms an explosive compound. It is, however, difficult of preparation: 40 parts of acetate of soda, 60 parts of quicklime, 40 parts of hydrate of potass are mixed together and placed in a retort. The lime merely prevents the hydrate of potass attacking the glass. The reaction consists of the conversion of the acetic acid, by the aid of the elements of HO, into carbonic acid, CO_2, and light carburetted hydrogen, CH_2.

$$\left.\begin{array}{r}\text{Acetic acid, } C_4H_3O_3\\ HO\\ \hline C_4H_4O_4\end{array}\right\} = \begin{array}{l}\text{2 eq. : } C_2O_4\\ C_2H_4 = 2CH_2\\ \hline C_4H_4O_4.\end{array}$$

This gas is colourless, invisible, inodorous; does not support combustion or respiration. It burns with a white flame, similar to olefiant gas, but less brilliant.

Coal Gas.—Coal gas is prepared by exposing coal to a red heat in iron retorts. The variety of coal known as cannel coal is usually employed. As the coal becomes heated a variety of complex gaseous products are given off. As these products issue from the retorts they are conducted into the hydraulic main (a large horizontal cistern half filled with water). The gas passes from the hydraulic main into the refrigerator and from the refrigerator into the purifier, and from the purifier into a large metallic gasometer—from which the distribution pipes proceed.

In the hydraulic main the gas is partially cooled and a portion of the ammonia absorbed. In the refrigerator the gas is cooled and the tar and ammoniacal liquor deposited. In the purifier, which is a large vessel containing slaked lime mixed with water and kept in agitation, the gas parts with its sulphuretted hydrogen and carbonic acid.

Ordinary coal gas, although subjected to this purification, usually contains a variable amount of noxious substances.

The compounds of carbon and hydrogen are the only valuable constituents of coal gas. It contains about fifty per cent. by volume of olefiant gas upon which its illuminating property chiefly depends; from twelve to twenty per cent. of light carburetted hydrogen (olefiant gas), with variable quantities of nitrogen, hydrogen, carbonic oxide, sulphuretted hydrogen, volatile oils, and prussic acid.

The Davy Lamp.—A mixture of fire-damp and air forms an explosive compound. Common explosions of gas are analogous to those which occasionally take place in coal mines. To prevent these accidents, a lamp was invented by Sir H. Davy, known as the Davy lamp.

Fire-damp alone is not explosive: but when mixed with from eight to ten volumes of air it is highly dangerous. With three or fifteen volumes of air it is not explosive. The small quantity of oxygen in the one case is not sufficient to burn the gas; in the other the large quantity of air conducts away the heat and lowers the temperature below the point necessary to ignite the gas. It is on this principle that the Davy lamp is constructed. Fire-damp does not take fire unless raised to a white heat. When the fire-damp issues from a seam of coal, meeting the white flame of a lamp or candle, explosion immediately takes place. The safety lamp consists of an ordinary oil lamp, surrounded by a cylinder of wire gauze, the apertures of the gauze not being less than the 400th of an inch in diameter. When this lamp is carried into an explosive mixture, no explosion takes place. The gases traverse the apertures of the gauze, and burn inside the lamp; but the flame does not pass outside to kindle the atmosphere of gas which surrounds the lamp. The lamp makes explosion impossible, not by preventing the gas entering inside, but by preventing the flame reaching the gases outside. In traversing the wire gauze the temperature is lowered, heat is conducted away, and the temperature falls below the point necessary to kindle the gases. These effects may be studied by suspending a Davy lamp in a jar, and gradually admitting a stream of coal gas from below. The flame of the lamp increases in length, in the interior of the gauze cylinder, the gas and air burn, and the flame of the lamp is extinguished; but as the atmosphere, in the bell jar becomes purer, the lamp again rekindles. These phenomena are so remarkable that the lamp becomes a true indicator of the state of the atmosphere in the mine; and this is the proper and true use of the Davy lamp.

Structure of Flame.—The flame of a candle is not as it appears, a solid cone of flame. If a piece of wire gauze be pressed from the top over the flame of a candle or common lamp, there appears in the centre a dark spot surrounded by a luminous ring. This dark spot is the combustible hydrogen and carbon of the tallow. These gases cannot burn until they pass to the surface, to obtain oxygen from the atmosphere, when they are converted into carbonic acid and water. The hydrogen, as we have seen, burns with a blue flame, with not much illuminating power, but with great heat. The oxygen of the air has a greater affinity for hydrogen than for carbon. The hydrogen burns first, and this is the blue flame around the central dark spot; the carbon now liberated from the hydrogen passes upward, and, being intensely heated by the combustion of the oxygen and hydrogen, burns with a yellow light; this is the luminous part of the flame, and is due to the solid white hot particles of carbon previously contained in combination with hydrogen in the tallow. The conical form of the flame is owing to the ascent of heated currents of air. A tall glass chimney is placed over lamps, so as to increase the upward current of air and supply a larger amount of oxygen. It is remarkable that the heat evolved by combustion is due to the quantity of oxygen consumed, and not the amount of combustible matter with which it unites, as the following Table illustrates :—

1 lb. of wood charcoal raised 78 lbs. of water from 32° to 212°.
1 lb. of alcohol raised 68 lbs. of water from 32° to 212°.
1 lb. of oil or tallow raised 90 lbs. of water from 32° to 112°.
1 lb. of hydrogen raised 236 lbs. of water from 32° to 212°.

The quantity of oxygen consumed under these circumstances is very different, while the weight of the combustible body was in each case one pound.

If a jet of air be introduced into the flame of a candle or lamp, by means of a blowpipe, a double combustion takes place. Two long pointed flames are observed, the outer one being of a yellowish colour, and the inner one blue. The space between these flames is filled with combustible matter at a very high temperature ; this is called the deoxidising or reducing part of the flame. If the oxide of any metal be placed in this part of the flame, fixed on the end of a fine platinum wire, or supported on a piece of charcoal, it is reduced to a metallic state.

Observations of great importance are often made with the blowpipe. The skilful use of this instrument can only be obtained by constant practice.

SILICON.

Symbol, Si; *Combining Equivalent*, 21·35.

This element is sometimes called silicium. It does not occur in a free state, but in union with oxygen it forms silica, SiO_3. It enters largely into the composition of flints, sand, quartz, and other mineral substances, forming it is thought a sixth part of the surface of the earth. Silicon is a brown powder, difficult to produce, and is of no importance, except to the scientific chemist. Silica, SiO_3, is the only known oxide. When nearly pure, as in rock crystal, it is colourless and transparent. It may be prepared by mixing equal parts of fluor spar, CaF, powdered glass, and sulphuric acid together in a flask. A glass tube, at least half an inch in diameter, should proceed from the flask and dip three quarters of an inch into mercury. The jar containing the mercury is now nearly filled with water; a gentle heat is applied to the flask, when the silica, in a gelatinous condition, separates and rises through the mercury to the surface of the water. The decomposition is rather complex. The first effect is the disengagement of hydrofluoric acid, and in contact with the silica of the glass, undergoes decomposition, forming water and the terfluoride of silicon, which escapes from the flask as a gas; but when it comes in contact with water, after it has passed through the mercury, it is decomposed into silica and hydrofluosilicic acid.

Decomposition in the flask:—

$$3HF + SiO_3 = 3HO + SiF_3.$$

Decomposition in contact with the water, after passing through the mercury:—

$$3SiF_3 + 3HO = SiO_3 + 3HF . SiF_3.$$

The silica is collected on a cloth filter, dried, and raised to a red heat to expel water.

Silica is an acid. It neutralises many bases, forming salts, known as silicates. Glass, porcelain, and pottery, are silicates of potash, alumina and other bases. Some silicates are soluble in water: if the bases with which the acid has united be in excess, the salts are soluble; but if the acid be in excess, then they are insoluble. Silica in its crystallised state, coloured with some metallic oxide, forms precious stones, such as blood-stone, agate, jasper, onyx, cornelian, and opal. Silica exists abundantly in many plants. It gives stiffness to the straw of wheat, barley, and oats.

BORON.

Symbol, B; *Equivalent*, 10·90.

This element is analogous to silicon. It may be obtained by heating boracic acid with potassium in a platina crucible. It is a brown powder, and, when heated, burns, forming boracic acid, BO_3. It is found in the lagoons of Tuscany, in combination with oxygen, as boracic acid, BO_3, from which source it is chiefly derived. This acid is also easily prepared by decomposing a hot solution of borax with sulphuric acid. The biborate of soda, or borax, is employed for soldering metals, forming the glazes for porcelain, and as a flux for examining metallic compounds before the blowpipe. A loop made of platinum wire is strongly heated and dipped into borax powder: a series of experiments can now be made. A salt of manganese heated in the outer flame of the blowpipe,—with the borax on the wire gives a violet-coloured bead; cobalt, blue; chromium, green; copper, red.

HYDROGEN.

Symbol, H; *Combining Equivalent*, 1.

Hydrogen, at all known temperatures, is a gas, colourless, without odour, about sixteen times lighter than common air, and highly combustible. This gas exists largely in the organic and inorganic kingdom; it is a constituent of all substances used either for light or illumination; it is a constituent of various acids, hydrochloric, HCl, hydrobromic, HB, hydriodic, HI, hydrosulphuric, HS. In combination with animal or vegetable substances it is found chiefly as water and ammonia.

Preparation. — This gas is always prepared by deoxidising water, either by the decomposition of water by voltaic electricity, as explained under oxygen, or by adding small pieces of granulated iron or zinc to one part by measure of sulphuric acid and four or five parts of water. Neither iron nor zinc combines directly with sulphuric acid, but the oxides of these metals combine easily. Decomposition: —

$$Zn + HO + SO_3 = ZnO,SO_3 + H \text{ free.}$$

Or hydrochloric may be used instead of sulphuric acid; then —

$$HCl + Zn = ZnCl + H.$$

By passing steam through a red-hot iron tube, the oxygen combines with the iron, and the hydrogen escapes as a gas, thus: —

$$\text{Steam, } HO + Fe = FeO + H.$$

By passing a small piece of sodium or potassium up into an inverted tube, filled with water and standing on the pneumatic trough:—

Process, $Na + HO = NaO + H$.

The sodium or potassium dissolves in the water, and the hydrogen collects in the upper part of the tube. Hydrogen from its extreme lightness, was formerly used for filling balloons; its use is now superseded by employing coal gas, and larger balloons. The property of diffusion, which is common, in a greater or less degree, to all gases and vapours, is seen best in the following experiment with hydrogen and carbonic-acid gas, two bodies of very different specific gravities. If two bottles, the one containing hydrogen, and the other carbonic acid, be joined together with a glass tube about a foot long and $\frac{3}{8}$ of an inch in diameter, and the hydrogen placed in the upper bottle and the carbonic acid in the lower, we shall find after a few hours, that the carbonic acid, contrary to its gravity, and the hydrogen, have uniformly mixed throughout. This diffusive power of gases is of great importance in the economy of the universe. Were it not for this diffusion we should have the different gases and vapours arranging themselves on the earth's surface according to their respective weights.

This diffusive power of hydrogen renders it impossible to keep it for any length of time in bladders, or the best constructed gas holders, and when mixed with air from any leaking, the mixture becomes explosive. When gases are much lighter or heavier than common air, they may be collected or examined without the aid of the pneumatic trough. A bottle or jar may be filled by inverting it over the extremity of an upright tube, from which the gas is issuing;

the tube should pass to the end of the bottle. When the gas is heavier, as chlorine, or carbonic acid, the reverse process must be adopted. The mouth of the jar to receive the gas should not be much larger than the tube from which the gas is delivered, if so, it should be covered with a piece of cardboard.

Hydrogen is not a supporter of combustion: if a lighted taper be plunged into a jar of hydrogen, it is extinguished; the hydrogen burns only at the top of the jar, where it comes in contact with the oxygen of the air.

Hydrogen burns with a pale blue flame without much light, but intense heat. The temperature may be increased by mixing the hydrogen and oxygen in the proportion of two volumes of the former to one of the latter, and burning it from a safety jet. This is called the oxyhydrogen blowpipe. It is, perhaps, best to use two separate gas holders for this experiment. When hydrogen is mixed with common air or pure oxygen, it is highly explosive, great care should be taken to allow the first portions of gas to escape before applying a light. Soap bubbles blown with hydrogen by means of a common pipe, attached to a bladder provided with a stop-cock to regulate the emission of the gas, rise in the air with considerable velocity: if a lighted taper be applied to the bubble as it ascends, a slight explosion takes place. If a mixture of one volume of oxygen and two volumes of hydrogen be taken, and soap bubbles blown on a small cheese plate, so that a good head of bubbles is formed and a light applied, the explosion is very violent: a small bladder filled with a mixture in the same proportion, may be exploded by pricking a hole with a pin and applying a light. If a jet of hydrogen is burned in a dry glass jar, the sides become covered with moisture, owing to the formation of water by the union of the inflamed hydrogen with the oxygen of the air. If a tube of glass about one inch in diameter, and two feet long, be slowly passed downwards over a burning jet of hydrogen, a musical note is produced. If a piece of spongy platinum be brought near a jet of hydrogen, it is at once kindled, and the platinum becomes red hot. It is thought that the gases are so condensed in this experiment, that they are brought within the sphere of attraction necessary for combination. The chemical relations of hydrogen assign to it the character of a metal, it has never been solidified, and its difference in physical properties to the other metals will probably disappear as the subject is more investigated.

There are two compounds of hydrogen and oxygen; water, HO, and the binoxide of hydrogen, HO_2.

The composition of water was first discovered by Mr. Cavendish, but its exact composition has occupied the attention of chemists for some years. There are two methods of determining the composition of any compound body; these are, analysis, by which the compound is resolved into its constituents, and synthesis, by which the elements or constituents are made to unite to produce the compound. The most elegant illustration of the analysis of pure water is its decomposition by voltaic electricity, alluded to under oxygen. In this experiment it appears that the combining volume of hydrogen is twice that of oxygen, and the combining proportion by weight, as 8 of oxygen to 1 of hydrogen.

The deutoxide of hydrogen, HO_2, or oxygenated water, is difficult of preparation. It has no use, and is only regarded as a chemical curiosity.

Water, when pure, is colourless, except when viewed in a large mass, when it appears of a beautiful blue colour. It is inodorous, tasteless, and transparent. It attains its maximum density at 40°; it boils under ordinary atmospheric pressure at 212°. It evaporates at all temperatures. It is 815 times heavier than air, and an imperial gallon weighs 10lb. avoirdupois. It is compressible in a very slight degree.

Water never occurs in nature perfectly pure. Filtration will remove suspended impurities, and distillation will free the liquid from those dissolved. Water unites with acids and bases, forming hydrates, this combination is sometimes very energetic, producing great heat, as in the slaking of lime, or by mixing sulphuric acid with water. The solvent properties of water are far greater than any other liquid. Many of the salts are soluble in water, increasing with an increase of temperature, and depositing crystals as the solution cools. Common salt and the salts of lime are, however, an exception to this general rule.

Water dissolves most of the gases, although in very unequal quantities. Hydrogen, oxygen, and common air are absorbed to a small extent, while ammonia and hydrochloric acid are dissolved to a very great extent; between these extremes we have sulphuretted hydrogen, chlorine, carbonic acid, and many other gases. The colder the water the more gas is dissolved; and on boiling, these gases are liberated.

RÉSUMÉ.

The Law of Definite Proportion.—Those bodies which combine chemically, combine in fixed and definite proportions.

The Law of Multiple Proportion.—When one body is capable of uniting with another in several proportions, these proportions bear a simple relation to each other, as O, O_2, O_3, O_4, &c.

Equivalent numbers are an expression of the relationship by weight of the different bodies which enter into chemical combination.

OXYGEN.

Symbol, O; *Equivalent,* 8.

Discovered by Priestley in 1774, called dephlogisticated air, afterwards vital air.

Physical Properties.—Permanently gaseous, without colour, odour, or taste.

Density, 1·106.

Chemical Properties.—A simple body soluble in water. It is present in all cases of combustion. The term combustion is used to express that chemical reaction usually attended with disengagement of light and heat. The common cases of combustion in nature consist in a slow absorption of oxygen, in which the temperature does not rise sufficiently to cause any sensible evolution of either light or heat. All cases of combustion are attended with the formation of new combinations.

Preparation.—By the decomposition of the binoxide of manganese in an iron bottle raised to a red heat—

$$2MnO_2 = Mn_2O_3 + O.$$

Second method. Decompose by heat in a Florence flask the chlorate of potash. Equal parts of the binoxide of manganese should be mixed with the salt. The manganese may be disregarded in the decomposition, as it undergoes no change—

$$KO,ClO_5 = KCl + 6O.$$

Third method. Decompose by heat the red oxide of mercury—

$$HgO = Hg + O.$$

Fourth method. By distilling in a glass retort the binoxide of manganese with sulphuric acid. The oxygen in this case is not very pure—

$$MnO_2 + SO_3 = MnO,SO_3 + O.$$

Fifth method. By the voltaic decomposition of water.

SULPHUR.

Symbol, S; *Equivalent*, 16.

Physical Properties.—At ordinary temperatures sulphur is solid, of pale yellow colour, without taste, having a peculiar and characteristic odour. It is opaque, and exists both in an amorphous and crystalline state. It is a bad conductor of heat and electricity. At a temperature of 230° it melts into a thin liquid; at a temperature of 430° it becomes thick and black like treacle; and at 790° it again becomes fluid and volatilises. The condensed vapour forms sublimed sulphur.

Sulphur burns with a pale blue flame.

Specific Gravity.—2·07.

Chemical Properties.—A solid body, insoluble in water, but soluble in the sulphide of carbon and essential oils.

Sulphur burns with a blue flame, forming sulphurous acid, SO_2. Sulphur combines with carbon at a red heat, forming the sulphide of carbon.

Sulphur is found in volcanic districts, and is imported to this country chiefly from Tuscany and Naples.

Uses.—For the manufacture of powder, lucifer matches, sulphuric acid, and in medicine.

SULPHUROUS ACID.

Symbol, SO_2; *Equivalent,* 32.

Physical Properties.—A colourless gas, with sharp odour and acid taste. It liquefies under a pressure of two atmospheres.

Specific Gravity.—2·234.

Chemical Properties.—Composed of one equivalent of sulphur and two of oxygen, very soluble in water, and possesses a strong acid reaction on test-paper. It is not combustible or decomposible by heat, nor is it a supporter of combustion or respiration.

Preparation.—By burning sulphur in air or oxygen. By heating in a flask copper and mercury with sulphuric acid.

Decomposition:—

$$\left.\begin{matrix}Cu\\Hg\end{matrix}\right\} + 2SO_3,HO = \left\{\begin{matrix}CuO,SO_3\\HgO,SO_3\end{matrix}\right. + 2HO + SO_2.$$

In preparing the hydrated sulphurous acid, carbon is substituted for copper.

Uses.—As a bleaching agent.

SULPHURIC ACID.

Symbol, SO_3; *Equivalent*, 40.

Known for a long time as oil of vitriol, analysed by Lavoisier.

Physical Properties.—Sulphuric acid in its anhydrous state is a white, solid, crystalline body. The hydrated acid is a colourless liquid with an oily appearance, without odour, and powerfully acid.

Specific Gravity of Sulphuric Acid.—As HO,SO_3, 1·843.

Chemical Properties.—Composed of one equivalent of sulphur to three of oxygen. The ordinary acid has the following composition, $2HO,SO_3$. It has a great affinity for water, and is used frequently as a desiccating agent. It carbonises organic substances, combining with the elements of water, and liberating the carbon. It is a powerful acid.

At a red heat it yields up an equivalent of oxygen, becoming sulphurous acid SO_2.

The nitrate of baryta will give a white precipitate on any solution containing sulphuric acid.

Preparation of Sulphuric Acid.—Fumes of the binoxide of nitrogen, NO_2, and sulphurous acid, SO_2, pass simultaneously into a leaden chamber, into which a jet of stem plays. The air in the chamber converts the binoxide of nitrogen into hyponitric acid, NO_4.

Two equivalents of sulphurous acid act on one of hyponitric acid, NO_4, reducing it to the binoxide of nitrogen, NO_2, which takes two equivalents of oxygen from the air in the chamber, converting it again into hyponitric acid, to be acted on by two more equivalents of sulphurous acid:—

$$O_2 + HO + NO_4 = 2SO_3 + HO + NO_2.$$

The acid is afterwards reduced to its proper strength by concentration.

Nordhausen acid.—This acid is prepared from the sulphate of iron and green vitriol by distillation. The vapours of sulphuric acid pass over, which are dissolved in water.

The anhydrous acid may be prepared by distilling the Nordhausen acid, and surrounding the receiver with a freezing mixture.

Sulphuric acid has a very extended application in the arts; it is used in fancy dyeing, and the manufacture of many chemical products.

SULPHURETTED HYDROGEN—HYDROSULPHURIC ACID.

Symbol, HS; *Equivalent*, 17.

Discovered by Scheele in 1776.

Physical Properties.—A colourless gas, having a strong, disagreeable odour, like rotten eggs. It liquefies under a pressure of 16 atmospheres.

Specific Gravity.—1·91.

Chemical Properties.—It is composed of one equivalent of hydrogen, and one of sulphur, HS. Very soluble in water, in which state it is usually employed in the laboratory. On exposure to air, the sulphur is precipitated. Sulphuretted hydrogen is a feeble acid; it is decomposed with sulphurous acid into water and sulphur:—

$$2HS + SO_2 = 2HO + 3S.$$

It is also decomposed by chlorine into hydrochloric acid and sulphur:—

$$Cl + HS = HCl + S.$$

The inhalation of this gas is very injurious, and chlorine is employed to decompose it. It frequently exists as a natural exhalation from drains and decomposing matter.

The gas decolorises some substances.

Preparation. — By decomposing the protosulphide of iron with sulphuric acid: —

$$FeS + HO,SO_3 = FeO,SO_3 + HS.$$

By decomposing the tersulphide of antimony with hydrochloric acid: —

$$SbS_3 + 3HCl = SbCl_3 + 3HS.$$

Sulphuretted hydrogen has no use in the arts.

CHLORINE.

Symbol, Cl; *Equivalent*, 35·5.

Discovered in 1774 by Scheele, and named dephlogisticated marine acid, muriatic oxygen acid, oxymuriatic gas.

Properties. — A gas of greenish colour, having a strong suffocating odour, somewhat resembling seaweed. It liquefies under a pressure of five atmospheres.

Specific Gravity. — 2·47.

Chemical Properties. — A simple body, very soluble in water. Chlorine has a great affinity for hydrogen. Many of the hydrogen compounds are decomposed with this element. A mixture of equal volumes of chlorine and hydrogen explode with great violence on exposure to sunlight.

Moist chlorine destroys organic colouring matters. It is also used as a disinfectant. Chlorine has a feeble affinity for oxygen.

Preparation of Chlorine Gas. — Heat gently in a retort or flask the peroxide of manganese and hydrochloric acid. Let them be well mixed before applying the spirit lamp: —

$$2HCl + MnO_2 = 2HO + MnCl + Cl.$$

The gas may be collected by displacement. If a solution of the gas be required, it must be passed into Wolfe's apparatus. By gently heating in a flask common salt, sulphuric acid, and the peroxide of manganese: —

$$MnO_2 + NaCl + 2SO_3 = NaO,SO_3 + MnO,SO_3 + Cl.$$

Uses. — As a bleaching agent, and for disinfection.

HYDROCHLORIC ACID.

Symbol, HCl; *Equivalent*, 36·6.

Discovered in 1650, and called spirit of salt, marine acid, muriatic acid.

Physical Properties. — A colourless gas, having a sharp, acid odour, and fumes on exposure to air. It is liquefied under a pressure of 40 atmospheres.

Specific Gravity. — 1·245.

Chemical Properties. — A body composed of one equivalent of chlorine united to one of oxygen. It is very soluble in water. The ordinary acid is a solution of this gas ($HCl + 16HO$).

Preparation of Hydrochloric Acid Gas. — Heating common salt in a flask with concentrated sulphuric acid —

$$NaCl + SO_3HO = NaO,SO_3 + HCl.$$

Uses. — For the extraction of gelatine from bones, and for disinfectants.

AQUA REGIA.

This is a mixture of one volume of nitric acid with three volumes of hydrochloric acid. Aqua regia is used to dissolve gold:—

$$3NO_5 + 6HCl + 4Au = 3NO_3 + 6HO + 2Au_2Cl_3.$$

Bromine and Iodine are analogous in their chemical properties to Chlorine.

AZOTE.—NITROGEN.

Symbol, N; *Equivalent*, 14.

Discovered by Lavoisier in 1775.

Physical Properties.—Permanently gaseous, without colour, odour, or taste.

Density, 0·972.

Chemical Properties.—A simple body, slightly soluble in water. It is neither a supporter of combustion nor animal respiration, nor is it combustible. Under ordinary circumstances its properties are remarkably inert, but in combination it forms compounds of very energetic and powerful character.

Preparation.—By burning phosphorus in a confined portion of common air. This is best done in a bell gas-jar standing over water. The phosphorus combines with the oxygen of the air, forming phosphoric acid, PO_5, which dissolves after a short time in the water.

Second method. By passing a current of air through a tube of hard glass loosely filled with copper tur [illegible] e tube is raised to a red heat. Oxide [illegible] ned, CuO, and the nitrogen liberated.

ATMOSPHERIC AIR.

First analysed by Lavoisier in 1777.

Physical Properties.—Permanently gaseous, without colour, odour, or taste.

Density, compared with water, 0·0013.

Composition of air as determined by Lavoisier:—100 parts by weight contain—

23 of oxygen
77 of nitrogen
100

By volume—

79·19
20·81
100·00

The air contains vapour of water, carbonic acid, and traces of ammonia.

Air may be analysed by the process described in the second method, by allowing the nitrogen to pass into an exhausted glass globe. Previous to the application of heat the glass tube and copper turnings are carefully weighed, and also the glass globe. The increase of weight after the experiment gives the information.

A less exact method is to introduce a piece of phosphorus, by means of a wire, into a graduated glass tube standing over water. The quantity of air in the tube is easily determined, by holding the tube upright and noticing the point to which the water rises. The whole is now left for twenty-four hours. The oxygen combines with the phosphorus, forming phosphoric acid, which dissolves in the water. The phosphorus may be withdrawn and the residual gas read off, which is nitrogen.

Air may also be analysed with the eudiometer.

Protoxide of Nitrogen (Laughing Gas).

Symbol, NO ; *Equivalent*, 22.

Physical Properties.—Permanently gaseous, without colour, having a sweet taste.

Specific Gravity.— 1·525.

Chemical Properties.— It supports combustion almost equal to oxygen. Cold water dissolves nearly its own volume of the gas. Warm water must be used in the pneumatic trough. The most remarkable feature of the gas is its exhilarating effects on the animal system. When inspired the gas should be quite pure, and only inhaled for a short time. It is prepared by heating the nitrate of ammonia in a retort or flask.

Preparation.—By decomposing the nitrate of ammonia by heat:—

$$NH_4O,NO_5 = 2NO + 4HO.$$

It has no use in the arts.

Binoxide of Nitrogen.

Symbol, NO_2; *Equivalent*, 30.

Physical Properties.—Permanently gaseous, and without colour.

Specific Gravity.— 1·038.

Chemical Properties.—Composed of one equivalent of nitrogen to two of oxygen. Insoluble in water,

to test-paper.

In contact with air or oxygen it changes into deep red fumes of hyponitric acid, NO_4.

It is not combustible.

Preparation.—By the decomposition of nitric acid by copper in the cold:—

$$4NO_5 + 3Cu = 3CuO,NO_5 + NO_2.$$

This gas has no use.

NITRIC ACID.

Symbol, NO_5; *Equivalent*, 34.

Sometimes called aquafortis, spirits of nitre.

Physical Properties.—When anhydrous it is a white crystalline solid. In its hydrated state it is a liquid. When concentrated it gives off fumes on exposure to air, forming nitric acid.

Specific Gravity of the Concentrated Acid.—1·522.

Chemical Properties.—Nitric acid is composed of one equivalent of nitrogen and five of oxygen. In its hydrated state it contains from one to two equivalents of water. Ordinary nitric acid contains from three to four equivalents of water, $NO_5,4HO$.

It is a powerful acid, and combines with water with great energy. At a red heat it is decomposed.

Nitric acid stains the skin yellow.

The hydrated acid may be prepared by heating gently in a flask the nitrate of soda with sulphuric acid.

Decomposition:—

$$NaO,NO_5 + 2HO,SO_3 = HO,NO_5 + NaO,SO_3 + HO,SO_3.$$

Nitric acid is largely used in the arts for etching, dyeing, &c.

AMMONIA.

Symbol, NH_3, HO; or NH_4O.

Sometimes called the volatile alkali spirits of hartshorn.

Physical Properties.—A colourless gas, having a suffocating odour. It liquefies under a pressure of seven atmospheres.

Specific Gravity.—0·597.

Chemical Properties.—Composed of one of nitrogen to three of hydrogen. Very soluble in water. It is a powerful base, turns vegetable reds blue, and combines with great energy with acids to form salts.

Ammonia is decomposed at a red heat. It is not combustible or a supporter of combustion; and to inhale the gas is very injurious,

Chlorine gas decomposes ammonia into the hydrochlorate of ammonia and nitrogen.

Preparation of Gaseous Ammonia.—

$$NH_4Cl + CaO = CaCl + HO + NH_3.$$

Liquid ammonia is prepared by dissolving the gas in water.

Ammonia is employed in medicine and dyeing. Its uses in nature are very numerous.

PHOSPHORUS.

Symbol, P; *Equivalent* 32.

This element was discovered by Brande, and made public in 1737.

Physical Properties.—A solid, generally of a pale straw colour, transparent, soft, and flexible. It is poisonous, and has a strong odour. In the dark it is luminous. There is another kind of phosphorus, called red or amorphous phosphorus, in which all the active properties of ordinary phosphorus are absent.

Specific Gravity.—1·77.

Chemical Properties.—A simple body, insoluble in water, but soluble in the bi-sulphide of carbon.

Phosphorus absorbs oxygen with great avidity. It combines with hydrogen in its nascent state.

Preparation.—By distilling at a red heat a mixture of charcoal and phosphate of lime (bone earth).

Phosphorus is extensively used in the manufacture of lucifer matches.

PHOSPHORIC ACID.

Symbol, PO_5; *Equivalent,* 72.

Physical Properties.—In its anhydrous state it is a white solid. There are three hydrates of this acid, all soluble in water:—

$$PO_5,HO\ ;\ PO_5,2HO\ ;\ PO_5,3HO.$$

Anhydrous phosphoric acid absorbs water with great avidity, passing gradually into its hydrated state.

Chemical Properties.—Phosphoric acid is a very energetic acid. It is decomposed with great difficulty.

Preparation.—By burning phosphorus in dry air or oxygen, white flakes of anhydrous phosphoric acid are formed, which readily dissolve in water.

By distilling phosphorus with nitric acid.

PHOSPHURETTED HYDROGEN.

Symbol, PH_3; *Equivalent*, 36.

Physical Properties.—A gas without colour, having a strong, disagreeable odour, like garlic.

Specific Gravity. — 1·13.

Chemical Properties. — Composed of one equivalent of phosphorus to three of hydrogen. Has no action on vegetable colours. Phosphuretted hydrogen inflames spontaneously in air, giving rise to phosphoric acid and water.

Preparation.—By heating a mixture of the hydrate of potash with phosphorus.

By decomposing water with the phosphide of calcium.

CARBON.

Symbol, C; *Equivalent*, 6.

Physical Properties. — Solid, very infusible, without taste, sometimes without colour; as the diamond. Generally of a black colour.

Density varies from 2 to 3·55.

It is found in nature crystallised and amorphous.

Carbon crystallised is a hard, dense body; a good conductor of heat and electricity.

Varieties.—The diamond, graphite, plumbago, coke, charcoal, bone-black, lampblack, anthracite.

Amorphous carbon is less dense, and a bad conductor of heat and electricity.

Chemical Properties. — A simple body, insoluble in water and other liquids. It has a great affinity for oxygen at high temperatures, and very combustible.

Preparation of Carbon. — Charcoal is prepared by burning wood with a limited supply of air.

Lampblack is prepared by burning oil or resinous substances. The lampblack collects in fine powder in the upper part of the chimney.

Coke is prepared by the combustion of coal in closed retorts, as in the manufacture of gas.

Bone charcoal is prepared by calcining bones in closed vessels.

Uses. — For the manufacture of pencils, India ink, black crayons, purifying water.

Animal charcoal is used as a decolorising agent in sugar-refineries.

Bisulphide of Carbon.

Symbol, CS_2; *Equivalent*, 38.

Discovered by Lampadius in 1796.

Physical Properties. — A thin, mobile, colourless liquid, with a most disagreeable odour, and very volatile.

Specific Gravity. — 1·293.

Chemical Properties. — Composed of one of carbon and two of sulphur. Without action on vegetable colours, insoluble in water, but very soluble in alcohol. It burns, giving rise to sulphurous and carbonic acid. It forms a detonating mixture with oxygen when inflamed.

Bisulphide of carbon is somewhat analogous to carbonic acid.

Uses. —Vulcanising india rubber and gutta percha, and killing moths.

LIGHT CARBURETTED HYDROGEN.

Symbol, C_2H_4; *Equivalent*, 16.

Known as inflammable air, marsh gas, fire-damp.

Physical Properties.—Permanently gaseous, without colour, but having a disagreeable odour.

Specific Gravity. — 0·556.

Chemical Properties. — It is composed of two equivalents of carbon and four of hydrogen, neutral in its action upon vegetable colours, and insoluble in water.

It is not a supporter of combustion or respiration.

It is combustible when mixed with common air; it detonates violently, forming carbonic acid and water.

Preparation.—Sixty parts of quicklime, forty parts of solid caustic potash, and forty parts of the crystallised acetate of soda, are strongly heated together in a retort. The reaction consists in the conversion of acetic acid by the elements of water into carbonic acid and light carburetted hydrogen. The affinity of potash for carbonic acid being the determining cause. The lime merely prevents the potash from fusing and attacking the glass.

Decomposition: —

$$\left.\begin{array}{r}\text{Acetic acid, } C_4H_3O_3 \\ HO \\ \hline C_4H_4O_4\end{array}\right\} = \begin{array}{l}2\text{ eq.} : C_2O_4 \\ \phantom{2\text{ eq.} : }C_2H_4 = 2CH_2 \\ \hline \phantom{2\text{ eq.} : }C_4H_4O_4.\end{array}$$

This gas has no use in the arts.

HEAVY CARBURETTED HYDROGEN.

Symbol, C_4H_4; *Equivalent*, 28.

Sometimes called olefiant gas.

Physical Properties. — At ordinary temperatures a gas, but not permanently gaseous. It is without colour, and has the odour of ether.

Specific Gravity. — 0·978.

Chemical Properties. — It is composed of four equivalents of carbon and four of hydrogen. Neutral to test paper, and insoluble in water. Mixed with air, it forms an explosive compound. When inflamed, it forms carbonic acid and water.

It is decomposed at a high temperature.

It is not a supporter of combustion or respiration.

When mixed in equal volumes with chlorine and a light introduced, an oily liquid is formed,—Dutch oil. If the chlorine be in excess there is an abundant deposit of carbon.

Preparation. — By heating in a flask one part by measure of alcohol, $C_4H_6O_2$, to five or six parts of sulphuric acid, SO_3,HO.

Decomposition: —

$$C_4H_6O_2 + SO_3,HO = SO_33HO + C_4H_4.$$

COAL GAS.

Obtained by the distillation of coal in iron retorts. It is a mixture of light carburetted hydrogen, heavy carburetted hydrogen, carbonic oxide, hydrogen, and other volatile products. The clearness of the light is owing to the presence of solid particles of carbon raised to white heat.

SILICON.

SILICIC ACID.

Symbol, SiO_3; *Equivalent*, 45·3.

Physical Properties. — A solid crystalline body, found in nature as quartz, agate, silex, flint, and sand. It is very refractory, scarcely yielding to the heat of the oxyhydrogen blowpipe.

Chemical Properties.—Composed of one of silicon to three of oxygen. It is insoluble in water and acids. Silicic acid unites with bases, forming silicates. It is decomposed by hydrofluoric acid.

Its uses in the arts are numerous.

Natural Silicates.—Felspar-clays.

Artificial Silicates.—Earthenware, glass, porcelain, &c.

HYDROGEN.

Symbol, H; *Equivalent*, 1.

Discovered by Cavendish in 1777, and called inflammable air.

Physical Properties.—Permanently gaseous, without colour, odour, or taste.

Density, 0·069.

Chemical Properties.—Simple body, slightly soluble in water. Has a great affinity for oxygen, and combines with that gas, forming an explosive mixture, which detonates violently on the application of a light.

It is combustible, but not a supporter of either combustion or respiration. It does not combine directly with nitrogen to form ammonia.

Preparation.—First method. Decomposition of cold water by zinc or iron in the presence of sulphuric acid—

$$Zn + SO_3,HO = ZnO,SO_3 + H,$$
$$Fe + SO_3,HO = FeO,SO_3 + H.$$

Second method. By passing steam through a gun barrel, loosely filled with iron turnings, and raised to a red heat—

$$4HO + 3Fe = Fe_3O_4 + 4H.$$

Uses.—Formerly employed to inflate balloons. The oxyhydrogen light of Drummond. The oxyhydrogen blowpipe.

PROTOXIDE OF HYDROGEN.—WATER.

Symbol, HO; *Equivalent*, 9.

Physical Properties.—A compound body, liquid at the common pressure and temperature. At 32° Fahr. it solidifies into ice. At 212° Fahr. it passes into steam.

Water is without colour, but in large masses it appears of a blue colour. It is without taste or smell.

Chemical Properties.—Composed, by weight, of 1 of hydrogen to 8 of oxygen; by volume, 2 of hydrogen to 1 of oxygen. The composition of water was discovered by Cavendish and Nicholson in 1800. The exact quantitative composition of water was determined by Gay-Lussac and Humboldt.

Analysis of Water.—First method. By the decomposition of steam when passed through a red-hot iron tube.

Second method. The decomposition of water by voltaic electricity.

Water is a solvent for most substances. Vegetable tinctures dissolve readily.

When water enters into composition with acids or salts they are called hydrated acids or salts; when destitute of water they are said to be anhydrous.

The purification of water is effected by distillation; such water is called distilled water, and is extensively used in the laboratory.

The uses of water are almost innumerable.

METALS.

POTASSIUM.

Symbol, K ; *Equivalent*, 39.

This metal was discovered by Sir H. Davy, in 1807, by exposing the hydrate of potash, KO,HO, to the action of a powerful voltaic battery between two plates of platinum.

Potassium occurs in nature in combination with silicic acid, SiO_3, in granite, trap, basalt, and other igneous rocks. These rocks crumble down by the action of the atmosphere, and the potash finds its way into the soil. Soils destitute of potash are uniformly barren. It is worthy of remark here, that the leaves and young parts of plants contain the most potash, while the trunks of trees contain scarcely any.

Potassium in combination occurs in most natural waters, and in the sea as the bromide of potassium, KB, the iodide of potassium, KI, and the chloride of potassium, KCl.

Potassium is prepared in the following manner:— A mixture of the carbonate of potash, KO,CO_2, and charcoal is prepared, and thoroughly dried in a covered iron pot. When quite cold it is reduced to powder, and mixed with $\frac{1}{10}$ of charcoal, in small lumps. The mixture is then transferred to a well hammered iron retort, or a bottle in which mercury is imported. The retort is placed in a furnace and the fire supplied with dry wood, so that the retort is enveloped in flame and kept at a uniform heat, approaching to whiteness. A copper receiver is divided by a partition, into one portion of which a tolerably large pipe from the retort dips into naphtha. The other side of the receiver is filled with ice. The fire is now raised in temperature. When decomposition of the carbonate of potash, KO,CO_2, by the charcoal takes place, carbonic oxide, CO, is abundantly disengaged, and large melted drops of potassium fall into the naphtha. Decomposition: —

$$KO,CO_2 + 2C = K + 3CO.$$

If the potassium is required very pure, it is redistilled with naphtha in an iron retort.

Potassium has a bluish white colour. It can be cut with a knife; at 32°, it is crystalline; at 70°, like treacle; at 150°, it liquefies. It has a sp. gr. of 0·865.

When potassium is exposed to oxygen, it loses its metallic lustre, and becomes potash, KO, which encloses the metal with a white powder. When thrown upon water, HO, the powder is dissolved off, and the pure metal comes in contact with the water, HO, and owing to its powerful affinities for oxygen, the water is decomposed into its elements, H and O. The high temperature caused by this combination is sufficient to ignite the hydrogen. If potassium is heated in dry air we obtain potash, KO, but if the air be moist, then hydrate of potash, KO,HO.

The hydrate of potash, KO,HO is a very important compound. It is always prepared by decomposing the carbonate of potash, KO,CO_2, with the hydrate of lime, CaO,HO, forming the carbonate of lime,

CaO,CO_2, which is insoluble, the hydrate of potash KO,HO remaining in solution :

Decomposition—

$$CaO,HO, + KO,CO_2 = KO,HO + CaO,CO_2.$$

The liquid containing the hydrate is boiled down, when all the water except one equivalent is evaporated. This furnishes the solid hydrate or caustic potash. It is raised in temperature till it melts, when it is run into cylindrical moulds about the size of lead pencils. It is kept in good stoppered bottles, owing to its powerful affinities for water, HO, and carbonic acid, CO_2.

The hydrate of potash, KO,HO, is soluble in water, highly alkaline, and neutralises most completely the most powerful acids. It dissolves the skin, and is used by medical men for that purpose. If a solution of the hydrate of potash KO,HO, be shaken with any of the fixed oils, we have soap.

Salts of potass, carbonate of potash, $KO,CO_2 + 2HO$, known as potashes or pearlash. It is brought to this country from America and Russia, where it is obtained from the ashes of burned wood (the potash being derived from the soil). The ashes are put into large barrels of water. After some time, the water, saturated with soluble carbonate of potash, KO,CO_2, is drawn off and evaporated to dryness in iron cauldrons. The potash is very impure, containing along with the carbonate, the silicate, sulphate, and chloride of potassium. The pure carbonate is obtained by calcining the tartrate of potash, digesting the mass with water, and evaporating the filtered liquid to dryness. Carbonic acid, CO_2, is too feeble an acid to completely neutralise the alkaline reaction of potash. It has therefore an alkaline taste and action on vegetable colours.

It is used in medicine and the manufacture of glass and soap.

The bicarbonate of potash, $KO,CO_2 + HO,CO_2$, is obtained by passing carbonic acid, CO_2, through a solution of the carbonate of potash, KO, CO_2, when you obtain a white crystalline precipitate of the bicarbonate of potash, $KO,CO_2 + HO,CO_2$. It is less soluble than the carbonate.

Nitrate of potash, KO, NO_5.—Nitre or saltpetre is brought to this country chiefly from the East Indies, where it appears, in some parts, like snow on the top of the ground. The nitre is collected, dissolved in water, to which a little carbonate of potash, KO,CO_2, is added, to throw down the lime. The liquid is evaporated, and we obtain the nitre of commerce. It is used for the manufacture of nitric acid, NO_5, and gunpowder. It has the property of preserving meat. In France, large quantities of nitrate of potash, KO,NO_5, are manufactured in the following manner: Heaps of old mortar and earth are collected and protected from the rain in a shed freely exposed to the air by having no sides, merely a roof to keep off the rain. These heaps are constantly turned and watered with putrid urine. When a good quantity of the salt has been obtained it is collected and mixed with water. The solution contains carbonate of potash, KO,CO_2, and nitrate of lime, CaO,NO_5. The filtered solution is then evaporated to dryness, and the salts purified by re-solution and crystallisation. The nitrate of potash, KO,NO_5, has no action on vegetable colours: it crystallises into six sided prisms, and is very soluble in water. When thrown on the metals in a state of fusion, oxidation occurs at the expense of the oxygen of the nitric acid, NO_5.

Gunpowder is a mixture of nitrate of potash,

KO,NO_5, sulphur, and charcoal, in the following proportions:—

Nitrate of potash, KO,NO_5 . .	75
Carbon, C	15
Sulphur, S	10
	100

The best charcoal for this purpose is manufactured from dogwood or elder. When this mixture explodes, the following rearrangement of the elements takes place: the oxygen of the nitrate of potash, KO,NO_5, goes to the carbon, forming carbonic acid, CO_2; the sulphur combines with the potassium, forming the sulphide of potassium, KS; and the nitrogen is set free. A volume of gas is formed occupying a space 1600 times greater than the powder, and this accounts for the violence of the explosion.

Sulphate of potash, KO,SO_3.—This salt is formed as the residue of the process for the manufacture of nitric acid, NO_5, from the nitrate of potash, KO,NO_5, in the presence of sulphuric acid, SO_3.

Decomposition—

$$KO,NO_5 + HO,SO_3 = HO,NO_5 + KO,SO_3.$$

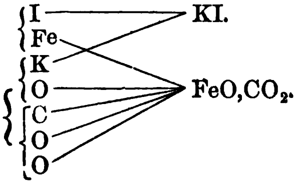
The sulphate of potash is soluble in water, and is used in medicine. The bisulphate of potash, $KO,2SO_3$ is used in calico printing.

Chlorate of potash, KO,ClO_5, is largely consumed for the manufacture of lucifer matches. It is prepared by passing a current of chlorine through a hot solution of caustic potash. Six equivalents of chlorine act on five equivalents of potash, KO, forming five equivalents of chloride of potassium, KCl, and one of chloric acid, ClO_5, which then unites with a sixth equivalent of potash, KO, forming five equivalents of chloride of potassium, KCl, and one of chlorate of potash, KO,ClO_5 :—

$$\begin{array}{l} \mathrm{Cl} \\ \mathrm{Cl,KO} \\ \mathrm{Cl,KO} \\ \mathrm{Cl,KO} \\ \mathrm{Cl,KO} \\ \mathrm{Cl,KO} \\ \hline \mathrm{5KCl + KO,ClO_5.} \end{array}$$

When this salt is heated, oxygen is abundantly disengaged, both from the acid and the base. The chlorate of potash, KO,ClO_5, deflagrates violently with combustible matter, explosion sometimes occurring by slight blows or touches. Sulphur and chlorate of potash, KO,ClO_5, explode with a loud report. It cannot be used for gunpowder for these reasons.

The iodide of potassium, KI, is prepared by adding iron filings to iodine water until the combination is complete and the solution colourless. The resulting iodide of iron, FeI, is filtered and decomposed with a pure solution of carbonate of potash, KO,CO_2, when the following decomposition takes place:—



$$\text{Or, } FeI + KO,CO_2 = KI + FeO,CO_2.$$

The iodide of potassium, KI, is separated by filtration and evaporation till the liquid is sufficiently concentrated to crystallise in cooling. It is very soluble in water, and crystallises in cubes. The bromide of potassium, KBr, may be prepared in the same way, substituting bromine for iodine.

All the salts of potash, KO, are without colour when not associated with a coloured metallic oxide or a coloured acid. They are all more or less soluble in water.

Tests for Potash.

If tartaric acid be added to a potash solution, a white crystalline precipitate of cream of tartar is thrown down. Agitation and a few drops of hydrochloric acid, HCl, promote the precipitation. Some time generally elapses before the appearance of the precipitate. Bichloride of platinum, in moderately dilute solutions of potash salts (on the addition of a little hydrochloric acid, HCl), gives a crystalline yellow precipitate, which is a double salt of bichloride of platinum and potassium. This compound and the cream of tartar are soluble in 60 parts of cold water. A little alcohol increases the delicacy of both tests.

Salts of potash colour the outer blowpipe flame purple or violet. This is only seen to advantage when the salts are pure. A little nitrate of potash thrown into a clear coal fire produces the same effect. A salt of potash, placed on a plate with a little alcohol and inflamed, produces a beautiful violet flame.

SODIUM.

Symbol, Na; *Equivalent,* 22·97.

This metal was obtained a few years after by Davy, by a similar method to that described under potassium. It is usually prepared by decomposing the carbonate of soda, NaO,CO_2, at a high temperature, with charcoal. Six parts of anhydrous soda are dissolved in a little hot water, and mixed with two parts of finely-powdered charcoal and one part of charcoal in lumps. The whole is then evaporated to dryness, transferred to an iron retort, and heated to whiteness. The tube of the retort passes into a receiver, containing rock-oil, and the whole operation is carried on the same as described under potassium. The process is easier than in the preparation of potassium, and more certain in its results. Sodium is a silver-white metal, soft at ordinary temperatures, melts at 195°, and oxidises in air, passing into soda, NaO. Its sp. gr. is 0·972.

Sodium thrown upon water decomposes it with great energy, the temperature seldom rising high enough to inflame the liberated hydrogen. In hot water it takes fire at once, burning with a yellow fláme, and giving rise to a solution of soda, NaO. Sodium resembles potassium in its properties.

Sodium occurs abundantly in nature, but not in its free state. Its most important native compound is the chloride of sodium, $NaCl$, which, as rock salt, occurs in immense beds in different parts of the earth, and sea water. Every natural water contains small quantities of this metal in combination. Small quantities of soda occur in most land plants, but largely in kelp and marine plants.

Soda, NaO, is obtained by exposing thin slices of the metal to dry air.

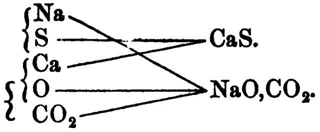
The hydrate of soda, NaO,HO, is obtained by decomposing a dilute solution of carbonate of soda, NaO,CO_2 with hydrate of lime, CaO,HO, the same as described under the hydrate of potassium, KO,HO.

The solid hydrate is white, highly alkaline, and a powerful solvent for animal matter. It is used in large quantities for making soap.

Salts of soda, carbonate of soda, $NaO,CO_2 + 10HO$. —This is one of the most important compounds of soda, from its consumption in glassmaking, bleaching, washing, baking, and medicine. It is obtained from various sources. One is from the ash of a plant which is cultivated on the shores of the Levant, called the salsola soda. It is produced in several places on the coast of Spain. This is the barilla of commerce, and contains about 18 per cent. of the carbonate of soda, NaO,CO_2. That made in Brittany from sea weeds, and called varec, contains about 5 to 6 per cent.

The native sesquicarbonate of soda, $2NaO,3CO_2$, is found as a natural product on the margins of certain lakes in Egypt, Africa, Mexico, and Hungary. The natron lakes of Egypt have yielded soda from a very remote period. By far the greater part of the carbonate of soda, NaO,CO_2, is manufactured from the sulphate of soda, NaO,SO_3, by the following method:—About 600 lbs. of chloride of sodium, $NaCl$, is placed upon the hearth of a well-heated reverberatory furnace, and an equal weight of sulphuric acid, SO_3, sp. gr. 1·6, is poured upon it from an opening in the roof. The sulphuric acid, SO_3, is thoroughly mixed with the salt; hydrochloric acid, HCl, is disengaged, which is either allowed to pass up the chimney or condensed. The salt by this process is converted into the sulphate of soda, NaO,SO_3, and hydrochloric acid, HCl.

This sulphate is reduced to powder and mixed with an equal weight of chalk, CaO,CO_2, and about half as much ground coal. Two cwt. of this mixture is thrown into a reverberatory furnace and heated to fusion, with constant stirring. When the decomposition is thought complete, the melted mass is raked into an iron trough, and allowed to cool. It is then broken up into little pieces, lixiviated with water, the solution evaporated to dryness, and the salt calcined with a little sawdust in a suitable furnace. This is the soda ash or British alkali of commerce, and, when good, contains about 50 per cent. of pure soda, NaO, partly as the carbonate of soda, NaO,CO_2, and partly as the hydrate of soda, NaO,HO, the remainder being the sulphate of soda, NaO,SO_3, chloride of sodium, $NaCl$, and traces of the sulphite and hyposulphite of soda. By dissolving this soda ash in hot water, and filtering the solution, the carbonate of soda, NaO,CO_2, is deposited in large crystals as the solution cools. The reaction which takes place in the calcination of the sulphate with coal dust is as follows:—The combustible matter of the coal converts the sulphate of soda, NaO,SO_3, into the sulphide of sodium, NaS; this sulphide of sodium, NaS, is then decomposed by the chalk, CaO,CO_2.



$$NaS + CaO,CO_2 = CaS + NaO,CO_2.$$

This process for the manufacture of the carbonate of soda, NaO,CO_2, was originally proposed by Leblanc, and although other methods have been proposed, this is found the most advantageous.

The ordinary crystals of carbonate of soda, NaO,CO_2, contain 10 equivalents of water, but by careful management they are sometimes obtained with seven, nine, fifteen, and sometimes with only one equivalent of water. When heated, they fuse in their water of crystallisation. The salt dissolves in two parts of cold water, and in less than its own weight of boiling water. The solution has a strong alkaline reaction and taste.

The bicarbonate of soda, $NaO,CO_2 + HO,CO_2$.—This salt is prepared by passing carbonic acid, CO_2, into a cold solution of the neutral carbonate, or by placing crystals of the bicarbonate of soda, NaO,CO_2, in an atmosphere of carbonic acid, CO_2. A solution of the bicarbonate of soda, $NaO,CO_2 + HO,CO_2$, is feebly alkaline, and much milder than the carbonate.

The sesquicarbonate of soda, $2NaO,3CO_2 + 4HO$, occurs on the banks of the soda lakes of Sokena in Africa. It is exported to this country under the name of trona. This is a bye product in some chemical manufactures.

Sulphate of soda, *Glauber's Salt*, product, $NaO,SO_3 + 10HO$.—It may be prepared by adding dilute sulphuric acid, SO_3, to a solution of the carbonate of soda, NaO,CO_2. This salt is slightly bitter and purgative. The mineral springs of Cheltenham contain this salt.

The bisulphate, $NaO,SO_3 + HO,SO_3 + 3HO$.—This salt is very soluble in water; has no acid reaction. When strongly heated it gives up an equivalent of sulphuric acid, SO_3, and becomes the sulphate of soda, NaO,SO_3.

The hyposulphite, NaO,S_2O_2, is largely used in photography. It is prepared by forming the neutral sulphite of soda by passing a stream of well-washed sulphurous acid, SO_2, into a strong solution of the carbonate of soda, NaO,CO_2; and digesting the solution with sulphur at a gentle heat for several days; then, by a careful evaporation at a moderate temperature, large and regular crystals of hyposulphite of soda, NaO,S_2O_2, are obtained, which are very soluble in water.

Nitrate of soda, NaO,NO_5.—*Cubic Nitre* occurs native in enormous quantities in Peru, where it forms regular beds to a very great extent. The nitrate of soda, NaO,NO_5, is used for making nitric acid, NO_5; but it cannot be used for gunpowder, as the salt burns too slowly and absorbs moisture.

The tribasic phosphate of soda, $2NaO,HO,PO_5 + 24HO$, is prepared by precipitating the phosphate of lime with the carbonate of soda.

A second salt, sometimes called the sub-phosphate, has the formula $3NaO,PO_5 + 24HO$. This is obtained by adding caustic soda to the preceding salt.

The tribasic phosphate of soda, and ammonia, and water, *Microcosmic Salt*, $NaO,NH_4O,HO,PO_5 + 8HO$. This salt is prepared by dissolving six parts of the common phosphate of soda in two of water, when the whole is liquefied. One part of chloride of ammonium, NH_4Cl, is added; when the chloride of sodium, $NaCl$, separates, which is removed by the filter; and from the properly concentrated solution, the microcosmic salt crystallises out.

Bibasic, phosphate of soda, $2NaO,PO_5 + 10HO$.

Monobasic, phosphate of soda, NaO,PO_5.

All the tribasic phosphates give a bright yellow precipitate with nitrate of silver, AgO,NO_5; the bibasic and monobasic give white precipitates with the same salt.

Biborate of Soda, *Borax*, $NaO,2BO_3 + 10HO$.—This salt occurs in the waters of certain lakes in Persia. The impure salt is imported to this country from the East Indies, under the name of tincal. A great quantity of this salt is now prepared from the native boracic acid, BO_3, of Tuscany, where it occurs, uncombined, in the hot springs of that country. Although this is an acid salt, it has an alkaline reaction to test-paper. It is used in the arts for soldering metals ; its action consists in rendering the surfaces of the metals to be joined metallic, by dissolving off the oxide. It sometimes enters into the composition of the glaze with which stone-ware is covered. The refined borax is used by the washerwomen of Holland and Belgium, to whiten linen and to soften water. About a handful is mixed with ten gallons of boiling water. The saving of soap is nearly one half, and the borax does not injure the linen.

Common Salt, chloride of sodium, $NaCl$.—A large quantity of salt is obtained from saline springs ; but our chief supply is from the mines of Cheshire. Rock salt is always too impure for use ; and if no natural brine-spring exists, one is formed by sinking a shaft into the rock salt, and introducing water. This water, when saturated, is pumped up, and evaporated in large iron pans. As the salt separates, it is removed by means of a wooden scoop, and pressed, while moist, into moulds, and then transferred to the drying stove. When large crystals are required, as for bay salt, the evaporation is conducted very slowly. Common salt generally contains a little chloride of magnesium, $MgCl$. The iodide of sodium, NaI, and the bromide of sodium, $NaBr$, are similar to the corresponding potassium compounds.

Tests.

All the soda salts are very soluble in water, except the antimoniate of soda. The presence of soda is generally determined by purely negative evidence, A solution of some salt of soda is prepared, the chloride of sodium, $NaCl$, or carbonate of soda, NaO,CO_2, may be taken, and five test-tubes half filled with the solution. The following reagents must be carefully added in the order given:—

Add the hydrosulphide of ammonium, NH_4HS; no precipitate.
Add carbonate of ammonia, NH_4O,CO_2; no precipitate.
The salt is a compound of potassium or sodium.
Add tartaric acid; no precipitate.
Add bichloride of platinum; no precipitate.
Quick lime, or caustic potash, and gently heated; no odour of ammonia.

Evaporate the solution to dryness, and observe the colour it gives to the outer flame of a blowpipe.

A bright yellow will indicate the presence of soda.

AMMONIUM.

Symbol, NH_4; *Equivalent*, 18.

In connection with potassium and sodium, ammonium may be conveniently studied.

Ammonia, NH_4O, is the oxide of a hypothetical body, called ammonium. The salts of ammonia correspond exactly with those of potash and soda. All attempts to isolate ammonium have failed, from its tendency to separate into ammonia, NH_3, and hydrogen. Ammonium plays the part of a metal in the composition of salts; it is capable of replacing sodium and potassium. The sal-ammoniacum of the Romans was the hydrochlorate or muriate of ammonia, and was obtained as a natural produce from a district in Libya. It was called by the alchemists spirits of hartshorn, because it was obtained by the distillation of deers' horns, in close vessels. It is sometimes called the volatile alkali.

Although ammonium has never been isolated, it is possible that it is not behind potassium and sodium in metallic lustre, and the ordinary physical characters of a metal.

If a small quantity of mercury be heated in a test-tube with a grain or two of potassium or sodium, and a gentle heat applied, violent combination ensues, attended with the production of light and heat. When cold, the fluid amalgam is put into a saucer, and covered with a strong solution of the hydrochlorate of ammonia. The mercury increases enormously in volume, and a pasty ammoniacal amalgam is formed. Left to itself, it quickly decomposes into mercury, ammonia, and water.

The great resemblance the ammoniacal salts bear to those of potash and soda, is the chief argument in favour of the existence of ammonium.

In the manufacture of coal gas, much ammonia is formed by the combination of the nitrogen and hydrogen of the coal. In the manufacture of bone charcoal ammonia is also produced; but in both cases it is very impure.

If a sal-ammoniac be reduced to powder, mixed with quick-lime, and heated in a retort, gaseous ammonia is abundantly evolved; and as water dissolves about 500 volumes of the gas, it must be collected over mercury, or by displacement. The vessel in which the gas is collected must be held above the orifice from which it is issuing, the same as if it were hydrogen. The following decomposition takes place:—

$$NH_3HCl + CaO = NH_3 + CaCl + HO.$$

Ammonia compounds are now largely manufactured from gas liquor, and the condensed products obtained from the distillation of bones and other animal refuse, as mentioned under bone charcoal. It was formerly obtained in Egypt from the soot of camel's dung.

The alkaline action of ammonia may be tried with turmeric, litmus paper, or purple cabbage water.

When ammonia comes in contact with one of the volatile acids, dense white fumes are formed. A test tube rinsed out with hydrochloric acid, HCl, and held over a vessel from which ammonia is issuing, white fumes and solid crystals of sal-ammoniac are formed by the union of the ammonia with the hydrochloric acid, HCl.

If gaseous ammonia, or ammonia water, be added to a solution of sulphate of copper, a splendid azure blue compound is formed. This colour is formed by a bluish precipitate, which is redissolved by excess of ammonia, and the blue colour produced afterwards.

Ammonia, when uncombined, may be readily recognised by any of the characters I have above described.

Salts of ammonia, Sal-ammoniac, NH_4Cl, is prepared from gas or bone liquor. The solution is treated with a slight excess of hydrochloric acid, HCl, by which the ammonia is neutralised, and the carbonate of ammonia, NH_4O,CO_2, and sulphide of ammonium, NH_4S, are decomposed, with evolution of carbonic acid, CO_2, and sulphuretted hydrogen, HS. The liquid is evaporated to dryness, and the salt carefully heated to expel the tarry matter with which it is associated. It is then purified by sublimation in iron vessels lined with clay and covered with domes of lead. This salt is sublimed without decomposition. Chloride of ammonium, NH_4Cl, forms double salts with the chloride of magnesium, MgCl, chloride of nickel, NiCl, chloride of cobalt, CoCl, chloride of zinc, ZnCl, chloride of copper, CuCl, and chloride of manganese, MnCl.

Sulphate of ammonia, NH_4O,SO_3+HO.—This salt may be prepared by neutralising the carbonate of ammonia with sulphuric acid; but it is largely manufactured for agricultural and commercial purposes by neutralising gas liquor with sulphuric acid, evaporating the solution, and purifying the product, which is impure sulphate of ammonia, $NH_4O,SO_2 + 10HO$. The salt $NH_4O,SO_3 + HO$ is soluble in two parts of cold water, and crystallises in long six-sided prisms, which lose an equivalent of water when heated. The sulphate of ammonia is entirely decomposed by ignition, and by long boiling with water the ammonia is volatilised, and the liquid rendered acid.

Carbonate of ammonia, NH_4O,CO_2. — The carbonates of ammonia are very numerous. The neutral anhydrous carbonate, NH_3CO_2, is prepared by the direct union of well dried and cooled carbonic acid, CO_2, and ammoniacal gas. These gases combine in the proportion of one volume of carbonic acid, CO_2, with two volumes of ammonia, NH_3. The compound is very volatile, has a pungent odour, is soluble in water, and condenses in white flakes. The carbonate of ammonia used in medicine is prepared by subliming a mixture of chloride of ammonium, NH_4Cl, and carbonate of lime, CaO,CO_2. This compound contains a less quantity of base than the neutral carbonate. Its composition varies, but in freshly-prepared specimens it approaches the sesquicarbonate, $2NH,HO,3CO_2$.

Sesquicarbonate of ammonia, $2NH_4O,CO_2$.—This salt is very unstable; it is decomposed by a slight increase of temperature. By exposure to air, it loses its odour, and becomes a soft, white powder, which is the bicarbonate of ammonia.

Bicarbonate of ammonia, $NH_4O,CO_2 + HO,CO_2$.—This salt, although volatile, is more permanent than the sesquicarbonate. The bicarbonate of ammonia, $NH_4O,CO_2 + HO,CO_2$, resembles, in every respect, the bicarbonate of potash, $KO,CO_2 + HO,CO_2$.

Nitrate of ammonia, NH_4O,NO_5, is prepared by adding dilute nitric acid, NO_5, to the carbonate of ammonia, NH_4O,CO, until the solution is neutralised by slow evaporation at a moderate temperature. The salt crystallises out in six-sided prisms, like the nitrate of potash, KO,NO_5, and deflagrates like that salt. When heated with combustible matter, it absorbs moisture, and is decomposed by heat into nitrous oxide and water.

The acetate and citrate of ammonia, NH_4O, are prepared by neutralising the carbonate with acetic or citric acid.

When ammonia, or its carbonate, is neutralised by hydrochloric acid, HCl, hydrosulphuric acid, HS, hydriodic acid, HI, and hydrobromic acid, HBr, the resulting salts are anhydrous. A dry acid is united to dry ammonia, and the salts have this formula: NH_3,HCl; NH_3,HS; NH_3,HI; NH_3,HBr.

The sulphide of ammonium may be prepared by distilling the chloride of ammonium, NH_4Cl, with the corresponding sulphide of potassium or sodium. The bi-sulphide of ammonia is a compound of great practical utility, $NH_4S + HS$. It is obtained by saturating a solution of ammonia with well-washed sulphuretted hydrogen until the solution is saturated.

By exposure to light, it becomes slightly yellow. It gives precipitates with most metallic solutions, several of which are very characteristic.

Tests.

All the salts of ammonia are more or less soluble. The acid tartrate and double chloride of ammonium and platinum being the least so. Ammonia, therefore, gives with the bichloride of platinum, $PtCl_2$, and with tartaric acid, $2HO,\bar{T}$, the same reactions as potash. The precipitates are similar in composition.

Ammonia may be recognised with certainty in the presence of all other bases by its being set free in a gaseous state, by the action of any of the caustic alkalies or alkaline earths (the hydrate of lime, CaO,HO, is the best for this purpose); and by its odour, and the white fumes produced when a glass rod moistened with hydrochloric acid, HCl (not fuming acid), is brought near the examined compound. There is no blowpipe test for ammonia.

BARIUM.

Symbol, Ba; *Equivalent,* 68

Barium was obtained by Sir H. Davy; it is a white, silver-looking metal, malleable, melts below a red heat, and slowly oxidises in air. It may be prepared by heating baryta, Ba,O, in a porcelain tube, through which the vapour of potassium is passed. The barium is then extracted by mercury, and the amalgam distilled in a small Bohemian glass retort.

There are two oxides of barium,—the protoxide, BaO, and the binoxide, BaO_2.

Baryta, or barytes, BaO. — The carbonate and sulphate of baryta occur in considerable quantities in some lead mines: from these salts baryta is obtained. If pure baryta is required, the crystallised nitrate of baryta is decomposed by heat in a porcelain crucible until red vapours of nitrous oxide cease to be disengaged. Baryta remains in the crucible as a grey spongy mass. If this residue be mixed with water, it forms a hydrate of baryta, BaO,HO. The hydrate of baryta has a great affinity for carbonic acid; and is soluble in water. A hot saturated solution deposits, on cooling, crystals having the following composition: $BaO,HO + 9HO$.

The hydrate of baryta is highly alkaline; and a valuable reagent. The smallest trace of carbonic acid instantly renders the solution turbid.

The binoxide, BaO_2, is prepared by raising to a red heat baryta, BaO, in a porcelain tube, through which a current of oxygen is passed.

Chloride of barium, $BaCl + HO$.—If the carbonate of baryta be dissolved in hydrochloric acid, and the solution filtered and slowly evaporated, crystals of the chloride of barium are deposited in the water of crystallisation.

Sulphate of baryta, BaO,SO_3, called heavy spar, is often found beautifully crystallised in lead mines. It may be produced when any soluble sulphate is added to a solution of any baryta salt. It is insoluble in water, or in acids. Boiling nitric acid dissolves a little, but the salt separates on cooling. Sulphate of baryta, BaO,SO_3, is used for painting, and adulterating white lead.

The Nitrate of baryta, BaO,NO_5, is prepared in the same way as the chloride, substituting nitric acid for hydrochloric, but the crystals are anhydrous. This salt is more soluble in pure water than in dilute nitric acid. A precipitate of the nitrate of baryta may be distinguished from the sulphate by its disappearance on the application of heat, or the addition of a large quantity of water.

Carbonate of baryta, *Witherite*, BaO,CO_3. This, like the sulphate, is found in some lead mines. It may be prepared by dissolving the barium chloride, $BaCl$, in water, and precipitating with the carbonate of soda or any other alkaline carbonate.

The hydrate of baryta is used in the laboratory to separate carbonic acid from gaseous mixtures; the nitrate and chloride to precipitate sulphuric acid.

Tests.

Sulphuric acid and the soluble sulphates precipitate immediately the sulphate of baryta, BaO,SO_3; insoluble in acids and alkalies.

Chromate of potash precipitates in neutral and alkaline solutions a pale yellow precipitate of the chromate of baryta; insoluble in acetic acid and alkalies.

Oxalic acid precipitates in concentrated solutions the white oxalate of baryta; soluble in acids. A few drops of ammonia promotes the formation of the precipitate.

STRONTIUM.

Symbol, Sr; *Equivalent*, 43·8.

This metal may be prepared by means like those described for barium, and is analogous to that metal in its physical and chemical properties. All the soluble salts of baryta are poisonous; this is not the case with strontia salts. It is found as a carbonate and sulphate in lead mines.

Protoxide of strontium, strontia, SrO. — This oxide may be prepared by decomposing the nitrate with heat, exactly similar to that described under baryta. The hydrate has a great affinity for carbonic acid.

The binoxide of strontia, SrO_2, may be prepared like the binoxide of baryta.

All the salts of strontia may be prepared by methods similar to those described under the salts of baryta.

The chloride of strontium, $SrCl$, is soluble in two parts of cold water. It is also soluble in alcohol, and when kindled, burns with a beautiful crimson flame.

The nitrate of strontia, SrO,NO_5, is used for the manufacture of red fire:—

	Red Fire.	Green Fire.
Nitrate of strontia (dry)	800 grains	450
Sulphur	225 „	150
Chlorate of potass	200 „	100
Lamp black	50 „	25

The strontia, sulphur, and lamp black must be well powdered and well mingled together, after which the chlorate of potass may be added in rather a coarse powder, and mixed with the other ingredients, without much friction. The red fire has sometimes ignited spontaneously.

Tests.

The soluble sulphates and sulphuric acid precipitate completely from its solutions the sulphate of strontia; the precipitate does not appear for some time, unless the solution is concentrated. The precipitate is insoluble in acids and alkalies.

Chromate of potash produces no precipitate, the chromate of strontia being soluble.

Oxalic acid precipitates the white oxalate of strontia in very dilute solutions; a little ammonia promotes the formation of the precipitate.

Salts of strontia, heated on a platinum wire in the inner flame of the blowpipe, give a crimson colour.

CALCIUM.

Symbol, Ca; *Equivalent*, 20.

This metal is obtained with difficulty. It has been seen by a very few persons. It is described as a silver-white metal, and has a great affinity for oxygen, passing into lime, CaO. The term calcareous is derived from the Latin word calx.

Calcium may be obtained by electrolysing in contact with mercury the hydrate of lime; the amalgam is then decomposed by distillation. It may also be obtained by strongly heating lime in an iron tube through which the vapour of potassium is passed. The reduced calcium is extracted by mercury, and the mixture distilled in a small green glass retort.

The *Protoxide of Calcium*, or *Quicklime*, is obtained by heating the carbonate of lime, CaO,CO_2. The black limestone of Kilkenny and Derbyshire is the best for this purpose. If lime be required absolutely pure, the nitrate of lime is precipitated with the carbonate of ammonia, and ignited to whiteness in a platinum crucible. Lime for ordinary purposes is obtained by calcining chalk in a kiln; a red heat continued for some hours is sufficient to liberate the whole of the carbonic acid. In the best lime kilns the process is carried on continuously by broken limestone and fuel being thrown in from the top and the lime raked out at intervals at the bottom. When the limestone contains silica, and the temperature is very great, the lime refuses to slake; the silicate of lime is formed, and the lime is said to have been overburned. Pure lime is quite infusible; at a high temperature it emits a pale light. When sprinkled with water it slakes and crumbles into a soft white powder, which is the hydrate of lime, CaO,HO. The equivalent of water may be expelled by raising it to a red heat, the hydrate returning to its original protoxide of calcium. This hydrate of lime is soluble in water; and the colder the water, which is an exception to the general rule, the more lime is taken up. A pint of water at 60° will dissolve about eleven grains; at 212°, about six.

Lime water may be prepared by agitating with water in a stoppered bottle the hydrate of lime; when the excess of lime falls to the bottom, the water may be poured off and a fresh quantity added. When lime is mixed with water the temperature is raised sufficiently to explode gunpowder and char wood; carts laden with lime have sometimes been set on fire by a shower of rain. If lime water be exposed to the atmosphere in an open vessel, it becomes covered with a thin film of the carbonate, which is due to the absorption of carbonic acid from the atmosphere. The hardening of mortar is partly owing to the gradual absorption of carbonic acid. Some cements, which are used for cementing the piers of bridges and other structures under water, contain alumina, silica, and oxide of iron. The cement has to be prepared with great care. When these cements are ground to powder and mixed with water, solidification ensues, and in this state it effectually resists the action of water. It is thought that a chemical combination takes place between the silica, oxide of iron, and alumina. In ordinary mortar no such combination takes place between the silica (sand) and the lime.

The beneficial use of lime for agricultural purposes is a matter of dispute; it is, however, useful in causing the rapid decomposition of organic matter in peaty soils. These soils are barren from the excess of vegetable matter. In clayey soils it acts beneficially by liberating the potass, which exists as an insoluble silicate in combination with alumina.

The *Binoxide of Calcium*, CaO_2, may be prepared by exposing lime to a red heat in a porcelain tube and a current of oxygen passed through the tube.

Salts of Lime, CaO,CO_2.—This is one of the most widely distributed compounds in the mineral world, forming whole mountains of immense thickness in almost every part of the world. It is the chief constituent of chalk, shells, and coral reefs. In nature the carbonate of lime is seldom pure; it generally contains clay, iron, and organic matter. But the most ancient of the crystalline limestones are destitute of organic remains, while the chalk is almost entirely made up of the exuviæ of once living beings. Limestone caverns are often lined with magnificent crystals of calcareous spar, which have been slowly deposited from a watery solution. These stalactite incrustations are often of great beauty. In pure water the carbonate of lime is scarcely soluble, but when the water contains free carbonic acid it is taken up. If lime water be added to a vessel containing carbonic acid, a milky-white precipitate is formed; but on agitation, if the carbonic acid be in excess, the turbidity disappears, and a clear, transparent solution of carbonate of lime is obtained. If the clear solution is boiled, the carbonic acid is expelled, and the carbonate of lime again precipitated. All natural waters contain carbonic acid. Lime in the condition of a carbonate is almost always present in river and spring water. In limestone districts it is often present to a very great extent. The hardness of water is usually owing to the presence of carbonate of lime. This hardness may be considerably removed by boiling; it is, however, most completely removed by the addition of lime water, when the carbonate of lime is precipitated. It is upon this principle that Professor Clark's process for softening water is based. The hardness of water containing carbonate of lime is called temporary, because it may be removed in a great degree by the methods mentioned. The use of such water for domestic purposes is attended with a great loss of soap; soap cannot act as a detergent until the whole of the lime is precipitated. Tea kettles and boilers in which such water is heated become lined with a thick incrustation. Dr. Ritterband has proposed a method of removing this deposit, which consists of throwing into the boiler a small quantity of sal ammoniac, when carbonate of ammonia is formed, which is volatilised by the high temperature, chloride of calcium remaining in solution. This method is not applicable to permanently hard waters. All the other salts of lime may be obtained from the carbonate by dissolving it in the various acids.

Sulphate of Lime, Gypsum, Alabaster, CaO,SO_3.—This salt is occasionally found, associated with rock salt, in an anhydrous and crystalline state. When regularly crystallised it is called selenite. If the chloride of calcium in solution be mixed with sulphuric acid, the sulphate of lime is precipitated; it is soluble in about 500 parts of cold water. The sulphate of lime is more soluble in water containing the nitrate of potash. Gypsum, or the hydrated sulphate, is largely employed in the arts for moulds, statues, and casts. The native hydrated sulphate, $CaO,SO_3 + 2HO$, if exposed to a temperature of 250° F. in an oven, is deprived of its water of crystallisation, and afterwards reduced to powder. When mixed with water it solidifies into a hard hydrate; but this solidification does not take place if the gypsum has been overheated. It is called by workmen plaster of Paris. Scagliola work, or artificial-coloured marble, is often prepared by inserting pieces of natural marble in the sulphate when the plaster is soft, and polishing the surface when the plaster has hardened. It is not possible to burn plaster when it has been once used, so as to make it set again. Sulphate of lime is a common impurity of spring water. When water contains this salt in solution it is said to be permanently hard, since this hardness cannot be removed.

Phosphates of Lime.—There are several compounds of lime with phosphoric acid, two of these are tribasic—

$$2CaO,HO,PO_5.$$
$$3CaO,PO_5.$$

These two salts are produced when the corresponding soda salts are added to a solution of the chloride of calcium. Bone earth consists of a combination of these salts. Bibasic and monobasic phosphates of lime exist. These are insoluble in water, but soluble in dilute acids.

Chloride of Lime, CaO,Cl.—This salt is extensively manufactured as bleaching powder. If moistened hydrate of lime be exposed to chlorine gas this compound is produced. The chlorine is at first supplied slowly, so as to prevent any great increase of temperature. It is soluble in 10 parts of water. The solution is highly alkaline and its bleaching properties are weak. When the hydrate of lime is suspended in cold water, and chlorine passed through the solution, the lime is gradually dissolved and the same bleaching compound is also produced. The hydrates and carbonates of the alkalies may, by similar methods, be made to absorb chlorine and form corresponding compounds, which are sold as disinfecting fluids.

The exact constitution of the bleaching compounds of lime is involved in some obscurity. It is thought that these salts contain hypochlorous acid, a body as remarkable for its bleaching properties as chlorine.

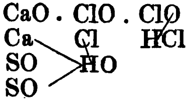
If good bleaching powder be distilled with dilute sulphuric or nitric acid, but not in sufficient quantity to decompose the salt, hypochlorous acid is obtained, hypochlorate of lime being formed simultaneously with the chloride of calcium. The following equation will illustrate the decomposition —

$$2CaO + 2Cl = CaCl + CaO,ClO.$$

If these salts be exposed to heat, their bleaching properties are entirely destroyed. If exposed in an open vessel they absorb carbonic acid and evolve chlorine.

Bleaching powder thus varies in value with its age, the care bestowed in its preparation, and its subsequent exclusion from the atmosphere.

The general outline of the method followed in bleaching is as follows: — The goods are soaked in a dilute solution of chloride of lime; they are then removed to a vat containing dilute sulphuric acid. The bleaching powder now undergoes decomposition by the acid.



The free chlorine disengaged in contact with the article to be bleached causes a complete destruction of the colouring matter. As the solutions of lime and sulphuric acid are very dilute, this process has to be repeated several times. If a pattern be stamped upon a coloured cloth with tartaric acid thickened with gum-water, and the cloth afterwards immersed in the lime-bath, no alteration of colour will take place except in those parts stamped with the acid. The acid acts, in this case, precisely the same as in the dilute sulphuric acid bath. The chlorine is disengaged, and the colour completely removed, in those parts stamped with the acid. We then obtain white patterns upon a coloured ground. If chloride of lime be mixed with water and placed in a saucer or basin, and a little acid added to the mixture, free chlorine is liberated; or cloths may be dipped in the solution and hung in rooms. One of these methods is usually adopted when chlorine is used as a disinfectant.

Chloride of Calcium, CaCl.—This salt is prepared by dissolving marble in hydrochloric acid. It is also a product in many chemical operations, and is used for drying gases. When mixed with snow or powdered ice it produces artificial cold. It is also used to determine the amount of hydrogen in organic bodies.

The *Sulphide of Calcium*, CaS, is obtained by reducing the sulphate of lime at a very high temperature with charcoal. This salt is slightly soluble in water. It is produced in the manufacture of carbonate of soda.

The *Fluoride of Calcium*, or *Fluor spar*, CaF, is the source from which hydrofluoric acid is prepared, and the other fluorides. It is found in beautiful cubic crystals, of various colours, in lead mines. It is decomposed by sulphuric acid.

The *Phosphide of Calcium*, CaP, is a brown-looking substance; it yields inflammable phosphuretted hydrogen when thrown on water. It is prepared by passing the vapour of phosphorus over fragments of lime heated to redness in a porcelain tube.

Tests.

Sulphuric acid and the alkaline sulphates give no precipitate, except in concentrated solutions of lime, and then only a partial precipitate, which completely dissolves in excess of water.

Chromate of potash gives no precipitate.

Oxalic acid throws down in dilute neutral solutions of lime the white oxalate of lime; a little ammonia renders the reaction more delicate.

MAGNESIUM.

Symbol, Mg; *Equivalent*, 12.

Magnesium is a white malleable metal, resembling zinc. It has no use in the arts, and is prepared by heating the chloride of magnesium with sodium in a test tube of German glass. Combination takes place with great energy and elevation of temperature; the chloride of sodium being formed, and metallic magnesium set free. When the tube is cold it is broken up and the contents put into cold water, by which the metal is separated from the salt.

Magnesia, MgO, is the only oxide of this metal. It may be obtained pure by raising to a red heat, in a porcelain crucible, the magnesia of the shops. It is a soft white powder, attracts moisture and carbonic acid from the atmosphere; unites slowly with water, forming a hydrate MgO,HO; it is slightly soluble in water. Its alkaline reaction is best observed by moistening a little, and placing it upon litmus paper. The most powerful acids are completely neutralised by magnesia.

Salts of Magnesia. — *Carbonate of Magnesia*, MgO,CO_2, occurs native, very much resembling calcareous spar.

If magnesia, MgO, be dissolved in water saturated with carbonic acid, and the solution left to evaporate, small prismatic crystals of the carbonate of magnesia are deposited with three equivalents of water. The carbonate sold in the shops is usually prepared by the addition of a hot solution of the carbonate of potass or soda to any soluble salt of magnesia. The precipitate, although usually called and sold as the carbonate, is not so. The salt is not constant in its composition. It usually has the following formula:—

$$4MgO,CO_2 + MgO,HO + 6HO.$$

Sulphate of Magnesia, Epsom salts, $MgO,SO_3 + 7HO$.—This salt is now largely manufactured by the action of sulphuric acid on magnesian limestone. It occurs in sea water and many mineral springs, is very soluble in water, has a disagreeable bitter taste, and purgative action. When heated, six equivalents of water pass off, the other equivalent being retained. Sulphate of magnesia forms double salts with the sulphate of potash and ammonia.

Phosphate of Magnesia, $2MgO,HO,PO_5 + 14HO$. —If a solution of the phosphate of magnesia and the phosphate of soda are mixed together, and allowed to remain some time, small prismatic crystals of the phosphate of magnesia separate. This salt exists in the grain of wheat, barley, and oats, and is found in considerable quantities in beer.

Phosphate of Magnesia and Ammonia.—*Microcosmic Salt*, $2MgO,NH_4O,PO_5 + 12HO$. — When any soluble phosphate is added to the sulphate of magnesia, a white precipitate slowly separates. If ammonia be added with the phosphate, and the solution concentrated, a crystalline precipitate appears at once; if the solution be dilute, the precipitate only appears after some time. This precipitate is the phosphate of magnesia and ammonia, $2MgO,NH_4O,PO_5 + 14HO$. This salt is slightly soluble in pure water. When heated, it is resolved into the bibasic phosphate. At a red heat, it fuses into a white enamel-looking mass, and is present in some forms of urinary calculus.

Silicates of Magnesia, MgO,SiO_3. *Steatite*, or *Soapstone*, chiefly found in Cornwall. *Meerschaum*, $MgO,SiO_3 + HO$, used for making pipe bowls. Chrysolite, $3MgO, SiO_3$, augite, hornblende, and jade, are composed of variable quantities of magnesia, silicic acid, alumina, and iron.

Chloride of Magnesia, Mg,Cl—This salt is obtained with some difficulty, owing to the obstinacy with which the last portions of water are retained. When magnesia is dissolved in hydrochloric acid, water and the chloride of magnesia are formed,

$$MgO + HCl = MgCl + HO.$$

This salt is very soluble, and the last portions of water are not evaporated from the solution without decomposition, the magnesia uniting with the oxygen and the hydrogen with the chlorine.

Tests.

The salts of magnesia are colourless.

The caustic alkalies and ammonia give a white gelatinous precipitate with magnesian salts, insoluble in excess, but soluble in chloride of ammonium.

Carbonates of soda and potash give a white precipitate. Carbonate of ammonia gives no precipitate in the cold.

Soluble phosphates give a white crystalline precipitate on the addition of a little ammonia.

MANGANESE.

Symbol, Mn; *Equivalent*, 27·6.

The metal manganese, or more properly the carbide, is prepared by heating the oxide with charcoal and borax. The process is rather difficult.

Manganese is only known as a chemical curiosity; it is put to no use in the arts. The metal is described as resembling cast iron, hard, brittle, and destitute of magnetic properties.

There are seven oxides of manganese, but two of these oxides are probably secondary compounds —

MnO, the protoxide,
Mn_2O_3, sesquioxide,
MnO_2, binoxide,
Mn_3O_4, red oxide $= MnO,Mn_2O_3$,
Mn_4O_7, varvicite $= Mn_2O_3.2MnO_2$,
MnO_3, manganic acid,
Mn_2O_7, permanganic acid.

The protoxide may be obtained in its anhydrous state by heating the carbonate of manganese in an atmosphere of hydrogen or steam; the carbonic acid is disengaged, and a green-coloured powder left behind, which is the anhydrous protoxide of manganese. It is best known in the form of the hydrate of the protoxide, and may be easily prepared by precipitating a salt of manganese with an alkali. The precipitate first appears white, then buff, then brown, by rapidly passing into higher states of oxidation. The protoxide is a powerful base, neutralising most completely the strongest acids.

Mn_2O_3.—This compound in union with water is found crystallised in some parts of Germany. If the hydrated protoxide be exposed to the air it passes into the sesquioxide. When oxygen is prepared by heating the binoxide in an iron retort, the residue is chiefly the sesquioxide, and is either of a black or brown colour, according to the circumstances under which it is produced. It is feebly basic. When heated with dilute sulphuric acid it dissolves; and on the addition of sulphate of potash deposits eight-sided crystals, having a similar constitution to common alum:—

$$Mn_2O_3,3SO_3 + KO,SO_3 + 24HO.$$

Binoxide.—This is the most common native compound of manganese. It is found amorphous and crystallised. It may be prepared artificially by adding a solution of bleaching powder to a salt of the protoxide of manganese. It is not obtained anhydrous by this method.

The binoxide is not soluble in water. It is of a black colour, and does not unite with acids to form salts. It is decomposed by hot hydrochloric and sulphuric acid, as in the preparation of chlorine and oxygen. This substance is largely used in the manufacture of bleaching powder.

Red oxide, Mn_3O_4, is found native. It is a reddish-looking substance, does not form salts. Borax and glass in a melted state dissolve this substance, and the glass acquires a beautiful amethyst colour.

Varvicite, $Mn_4O_7 = MnO + 2MnO_3$. — This compound resembles the binoxide, but is harder and more brilliant. It was first obtained from some manganese ores of Warwickshire.

The *Chloride of Manganese*, MnCl, may be prepared from the residue left in the retort after the preparation of chlorine from the binoxide of Mn and HCl. A solution of the chloride of manganese, has usually a pink colour, which deposits crystals of the same colour on evaporation. It dissolves readily in alcohol, with which it forms a crystallisable compound.

Sulphate of the Protoxide of Manganese, $MnO,SO_3 + 7HO$, is largely used in dyeing. It has a beautiful rose colour, and is prepared by raising to a red heat in a closed vessel the binoxide of manganese and coal, and dissolving out the impure oxide thus obtained with sulphuric acid. Towards the end of the process a little hydrochloric acid is added; the solution is evaporated to dryness, and again exposed to a red heat. Water is then added, which dissolves out the pure sulphate of manganese. The salt is used to produce a permanent brown dye. It forms, like alumina, a double salt with potash.

MnO,CO_2, is prepared by precipitating the protochloride of manganese with any alkaline carbonate. It is not soluble. When exposed to heat, it parts with its carbonic acid, and absorbs oxygen.

MnO_3.—If the binoxide of manganese be fused with an alkali, an additional quantity of oxygen is absorbed from the air, and a deep green mass is obtained, which is a salt of manganic acid. The presence of nitrate of potash facilitates the production of this acid. MnO_2, by combining with another equivalent of O, becomes MnO_3, which unites with the alkali, forming KO,MnO_3. It is of a green colour, soluble in water, and if evaporated in vacuo yields green crystals.

Permanganic Acid, Mn_2O_7.—If the manganate of potash, free from excess of the alkali, is put into a large quantity of water, it is decomposed into the hydrated binoxide of manganese, which falls to the bottom and forms a deep purple liquid, which contains the permanganate of potash. The changes of colour which accompany this decomposition have procured for the substance the name of mineral chameleon. When the potash is in excess the manganate is more stable, and the reaction is very slow.

The manganates and permanganates are decomposed by contact with organic matter.

Nearly all the salts of this metal are soluble in water, giving rose-coloured solutions. All the salts are soluble in hydrochloric acid.

Tests.

Hydrosulphuric acid, HS, produces no precipitate with a solution of the salts of manganese. NH_4S precipitates the metal completely as a flesh-coloured sulphide, soluble in strong acetic acid, and passing, by the absorption of oxygen, into a brownish black colour.

The fixed caustic alkalies precipitate the metal as a white hydrate, which becomes brown.

The alkaline carbonates precipitate the white carbonate of the oxide of manganese.

The carbonates of the alkaline earths do not decompose manganese salts, except by an increase of temperature.

If any compound of manganese be fused with carbonate of soda on a platinum foil, the green manganate of soda is produced. This is a very delicate test.

An amethyst bead is produced in the oxidising flame of the blowpipe when any compound of manganese is fused with borax, which disappears in the reducing flame.

ALUMINUM.

Symbol, Al; *Equivalent*, 13·69.

There is but one oxide of this metal, namely, alumina, Al_2O_3.

In combination with silica and potash, as the silicate of alumina and potash, $Al_2O_3,3SiO_3 + KO, SiO_3$, it is very abundant in nature, forming felspar, one of the constituents of granite, from which our various modifications of clay are derived. Clay consists chiefly of a silicate of alumina, $Al_2O_3,3SiO_3$. Pure clay, such as that employed for porcelain manufacture, is quite white: it is derived from the colourless felspar occurring in the granite rocks of Cornwall and Devonshire. Clay is sometimes red, owing to the presence of oxide of iron; sometimes blue or black, from the excess of vegetable matter.

The pure metal aluminum may be obtained by introducing into a platinum tube closed at one end the sesquichloride of alumina, Al_2Cl_3, with an equal quantity of potassium loosely wrapped up in platinum foil. The closed end of the tube is then heated, so as to sublime the chloride, bringing its vapours into contact with the melted potassium. The sesquichloride of alumina is then reduced. The metal is separated from the chloride by cold water. It has metallic lustre, it is malleable and of a tin-white colour. When heated in air or oxygen, it takes fire and returns to its original oxide, Alumina, Al_2O_3. If a solution of common alum be mixed with an excess of ammonia, a white gelatinous precipitate of the hydrate of alumina is thrown down. This precipitate is washed, dried, and ignited to whiteness. The alumina so obtained is a white, tasteless mass, scarcely affected by acids. It is fusible before the oxyhydrogen blowpipe. The ruby and sapphire are transparent varieties of crystallised alumina, and emery powder is a coarse variety of this oxide. Alumina is a very weak base, and its salts have frequently an acid reaction.

The *Sesquichloride of Alumina*, Al_2Cl_3, may be prepared from pure alumina (obtained by the method just described) mixed with charcoal and strongly calcined in a covered crucible. The contents of the crucible are then transferred to a porcelain tube heated to redness in a furnace; a current of chlorine passes through the tube when the sesquichloride of alumina condenses in the cold part of the apparatus. It is a crystalline substance and very soluble in water.

Sulphate of Alumina, $Al_2O_3,3SO_3 + 18HO$, is obtained by boiling sulphuric acid with clay, or by saturating dilute sulphuric acid with the hydrate of alumina, ($Al_2O_3,3HO$), and evaporating, when thin crystallised scales of the sulphate of alumina are obtained; it has an acid reaction, and when heated to redness is decomposed. The sulphate of alumina combined with soda, potash, and ammonia, forms double salts. All the preparations of alumina are obtained from common alum, which is a double salt of potash and alumina, $Al_2O_3,3SO_3 + KO,SO_3 + 24HO$. Alum is chiefly manufactured from a slaty kind of clay which contains the bisulphide of iron. This alum schist, as it is called, is first gently roasted, then moistened with water and exposed to the atmosphere. The sulphur of the iron is converted into sulphuric acid by the absorption of oxygen from the air. The sulphate of the protoxide of iron and the sulphate of alumina are produced. These salts are separated by lixiviation with water. The solution is concentrated by evaporation and mixed with the chloride of potassium which decomposes the iron salt forming the protochloride of iron and sulphate of potash. This sulphate of potash then combines with the sulphate of alumina to form alum. The chloride of iron which is very soluble in water, is separated from the alum by crystallisation.

By a repetition of this process the alum is purified. It is largely used in preparing skins, dyeing, and pigments. Alumina has a great attraction for the fibre of cloth and organic colouring matter. If a solution of common alum be filtered through a piece of calico, part of the alumina will be abstracted and retained by the cloth. If this cloth be then rinsed in an infusion of madder or indigo, a colour is produced, which is insoluble in water, and unaffected by soap. Substances which act in this manner are called mordants. The attraction of alumina for organic colouring matter may be also shown by adding to an infusion of litmus a solution of alum, and then a solution of soda, so as to precipitate the alum. As the precipitate falls, the colouring matter will be carried with it. Such precipitates are called lakes. Ammonia and soda may be made to replace the potash in sulphate of alumina and potash.

The *Silicate of Alumina*—$Al_2O_3,3SiO_3$,—is the next most important salt, since it enters into the composition of many crystallised substances. In the primary unstratified rocks, the silicate of alumina is very abundant. All clays are silicates of alumina, but their composition is very variable. By boiling with sulphuric acid the alumina is dissolved out, and the silica left behind. Clay containing the carbonate of lime is called marl. These may be easily recognised by their effervescing with hydrochloric acid.

Tests.

Alumina, in solution, can be distinguished without difficulty.

Caustic potash, or soda, added drop by drop, will occasion a white gelatinous precipitate, which is the hydrate of alumina, which is soluble in an excess of the reagent:—

$$Al_2O_3,3SO_3 + 3NaO,HO =$$
$$Al_2O_3,3HO + 3NaO,SO_3.$$

Ammonia produces a similar precipitate.

The alkaline carbonates and the carbonate of ammonia precipitate the hydrate of alumina, with the escape of carbonic acid. The precipitates are insoluble in excess.

IRON.

Symbol, Fe; *Equivalent*, 28.

This is one of the most important of the metals. The progress of commerce and civilisation is intimately connected with the uses and properties of iron. The pure metal is exceedingly rare, but its compounds are widely distributed throughout nature. It forms a great part of the colouring matter of many rocks, it is contained in plants, and forms an essential component of the blood. Pure iron is of a white colour, and crystallises in the form of a cube. The fibrous texture of the iron (upon which its strength mainly depends) is not natural to the metal; it is produced by hammering and rolling, and may always be observed when the metal has been attacked by rust or the application of an acid. Iron is the most tenacious of all metals; a wire $\frac{1}{38}$ of an inch in diameter will sustain a weight of 60 lbs.

The chief source of this metal is the carbonate of the protoxide. The most important and abundant source is the clay ironstone, which consists of the carbonate of the protoxide, which is usually associated with clay, the carbonates of lime and magnesia, silica, manganese, and coaly matter. This ore contains from 20 to 45 per cent. of iron. The iron ore is first roasted, to get rid of the water, carbonic acid, and carbonaceous matter. The roasted ore is then introduced into a blast-furnace. The ore is introduced with lime and coal at the top of the furnace, the two latter substances being employed as reducing agents. A very high temperature is kept up by forcing air in at the bottom of the furnace.

Formerly cold air was employed, but now the air is heated by passing it through hot pipes; the air by this means is raised to a temperature of 600° or 800°, and this is called the hot blast. This effects a great saving of fuel, and coal may be used instead of coke. The limestone introduced into the furnace separates the silicate of alumina from the iron, so that the coal can act upon the metal. The clay combining with the lime forms a glassy compound of the silicate of alumina and lime, which is called slag. The oxide of iron parts with its oxygen to the coal. As the iron melts it falls to the bottom of the furnace, uniting in its descent with the white hot carbon of the coal, forming a carbide of iron, which accumulates in a stratum at the bottom. The silicate of lime and alumina, having less specific gravity than the melted iron, floats like cream upon the surface. The iron and slag are drawn off at different openings at the bottom of the furnace. The iron is run into groves made of sand, and this constitutes the pig-iron of commerce. This iron is associated with carbon, and generally with silicon, manganese, magnesium, calcium, sulphur, and phosphorus. Cast iron is crystalline, brittle, and much more fusible than wrought iron.

Malleable wrought iron is prepared from cast iron, by a process called puddling. The object of this process is to deprive the iron of its carbon. The cast iron is laid on the floor of a reverberating furnace. The stream of burning gases is bent over upon the iron lying on the bed of the furnace. The workman, by means of a large iron rod like a boat-oar, stirs up the melted mass, so as to bring every part successively to the surface. The object of this is to oxidise the carbon of the iron by contact at a high temperature with oxygen. The carbonic oxide thus formed burns with its beautiful blue flame on the surface of the melted iron. Water is sometimes sprinkled on the melted mass, so as to supply oxygen to the carbon. The withdrawal of the carbon diminishes the fusibility of the metal. It first becomes thick and treacly, then like coarse sand. The temperature of the furnace is now increased; the iron again becomes about the consistence of treacle. The workman gathers the iron round his puddling-rod into large balls, which are exposed to hammering by steam hammers. All the liquid cast iron contained in these balls is squeezed out, and the particles of iron welded together in one mass. This mass is heated a second time and passed between large iron rollers. The fibrous structure of iron is mainly due to this process. Iron which is to sustain great tension is submitted, with slight modification, to this latter process several times. This iron is no longer crystalline; it is tough and fibrous in its texture, which may be seen by dissolving a piece of iron wire in dilute acid. By violent concussion this toughness and fibrous character disappears, and it again becomes crystalline and brittle. This malleable iron, although much purer than the cast iron, is never free from carbon and other elements.

The purest and best malleable iron is made into steel by placing alternately in an iron box bars of iron with charcoal powder, and exposing the whole to high red heat for forty or fifty hours. The iron has combined with a small portion of the carbon. It is now more fusible than malleable iron, but not so fusible as cast. It has also lost in a great degree its malleability. If fine steel is required, several bars are welded together under the steam hammer, which gives greater uniformity of composition.

Steel possesses great hardness and elasticity. If raised to a high temperature and plunged into water it becomes elastic, hard, and brittle; if it be slowly cooled it softens, loses its hardness and elasticity, and becomes very much like malleable iron. By altering the temperature and the rapidity of cooling every degree of hardness can be obtained. This is called tempering.

Iron does not rust or oxidise in dry air at the common temperature; heated to redness it becomes covered with the black oxide, the same compound as is produced when iron wire is burnt in oxygen. Iron at a red heat decomposes water, like potassium at the common temperature, liberating the hydrogen. Dilute sulphuric and hydrochloric acid dissolve iron freely. If magnetic iron be raised to a red heat it loses its magnetic properties.

There are four oxides of iron —

FeO, protoxide,
Fe_2O_3, sesquioxide, or peroxide,
Fe_3O_4, protosesquioxide, or black oxide,
Fe,O_3, ferric acid.

The *Protoxide*, FeO, has strongly marked basic properties; it neutralises acids most completely. It is scarcely known in its separate state, owing to its proneness to pass into the sesquioxide. If any salt of the protoxide be mixed with caustic alkali a bulky white precipitate of the hydrated protoxide is thrown down; this hydrate when exposed to air becomes first green, then reddish brown. The soluble salts of the protoxide of iron have usually a faint green colour.

The *Sesquioxide*, Fe_2O_3, occurs as a native compound, both amorphous and crystalline. The island of Elba has furnished some very fine specimens. The sesquioxide of iron in powder has a deep red colour, and is used for painting. It is prepared by calcining the sulphate of the protoxide. The colour varies according to the temperature to which the sulphate has been exposed. The sesquioxide is reduced at a high temperature in the presence of carbon or hydrogen, but in ordinary fires it undergoes no change. It is a feeble base, dissolves with difficulty in acids, and its salts have an acid reaction.

The *Black Oxide*, $FeO + Fe_2O_3$, is a natural product, known as the magnetic iron ore; it is one of the most important iron ores. This compound is the chief product of burning iron in oxygen. It does not form salts by union with any acid.

Ferric Acid, FeO_3.—This compound is rather unimportant.

Protochloride of Iron is prepared by passing well-dried HCl over red-hot malleable iron, or by dissolving iron in HCl. The solution when duly concentrated yields green crystals of the protochloride with four equivalents of water.

Sesquichloride, Fe_2Cl_3, is generally prepared by dissolving the sesquioxide in HCl. The solution, concentrated and evaporated, yields crystals of a red colour, which are very soluble in water and alcohol.

Protoiodide of Iron, FeI, is an important medicine. It is easily prepared by digesting iodine with water and metallic iron. The solution, which is of a green colour, yields on evaporation crystals of the protoiodide of iron.

There are several compounds of iron with sulphur, but only two are important. The *protosulphide*, which is a black, brittle-looking substance, from which hydrosulphuric acid is prepared. The protosulphide of iron, FeS (salts of iron), is obtained by throwing into a red-hot crucible two and a half parts of sulphur and four parts of iron filings. The *bisulphide*, FeS_2, is a natural product; it is formed in some cases by the gradual deoxidation of the sulphate of iron by organic matter. When heated, an intermediate sulphide is produced analogous to the black oxide.

Sulphate of the Protoxide, $FeO,SO_3 + 7HO$.—This salt is prepared by dissolving iron in dilute sulphuric acid. It is often prepared on a very large scale by contact with air and moisture with iron pyrites, which, by absorption of oxygen, furnishes the sulphate of iron. The salt is dissolved out by water, and the solution crystallised. It forms large green crystals.

Sulphate of the Sesquioxide, $Fe_2O_3,3SO_3$.—This salt yields compounds with the sulphate of potash and ammonia, analogous in their constitution to common alum.

Carbonate of Protoxide of Iron.—If any of the protosalts of iron be mixed with the alkaline carbonates, a white precipitate is obtained. This compound loses its carbonic acid when washed and dried, passing into a higher oxide. This compound occurs in nature as the clay ironstone, from which nearly all the British iron is manufactured. It is found in many mineral waters, being soluble in excess of carbonic acid. Water containing iron is known by the rusty matter it deposits by exposure to the air.

The salts of the protoxide of iron are colourless, but if concentrated have a pale greenish colour; when anhydrous they are white. They absorb oxygen, passing into the sesquioxide, which from neutral solutions is deposited as a yellow salt.

Tests.

HS forms only a slight precipitate of the black sulphide in neutral solutions; in acid solutions no precipitate is produced except in the acetate.

NH_4S throws down iron as a black precipitate, insoluble in excess. When any non-volatile organic acid is present a small quantity of the sulphide remains suspended, imparting to the solution a light green colour; on heating, this is deposited, and the liquid becomes of a bright yellow colour.

The caustic alkalies precipitate from salts of the protoxide of iron a white hydrate of the oxide, which on exposure to the air becomes a dirty green, then a reddish brown.

Ammonia will completely precipitate the salts of the protoxide of iron when no ammoniacal salts are present.

Alkaline carbonates precipitate the white carbonate of the oxide, soluble in the chloride of ammonium.

Carbonates of the alkaline earths give no precipitate.

Ferrocyanide of potassa gives a precipitate in solutions quite free from the sesquioxide and common air, a white precipitate of the ferrocyanide of iron and potassium, $K + 3Fe + Cy$. This precipitate assumes a bluish colour when the sesquioxide and air are present, and passes into Prussian blue.

Ferricyanide of potassa produces with salts of iron a blue precipitate.

These precipitates are insoluble in acids, but are easily decomposed by the alkalies.

All the soluble persalts of iron form, with water or hydrochloric acid, brownish yellow solutions.

Metallic iron, sulphurous acid, and hydrosulphuric acid reduce salts of the sesquioxide to salts of the protoxide. The hydrosulphuric acid renders the solution milky by the precipitation of sulphur —

$$Fe_2O_3,3SO_3 + Fe = 3(FeO,SO_3).$$
$$Fe_2O_3,3SO_3 + SO_2 = 2(FeO,SO_3) + 2SO_3.$$
$$Fe_2O_3,3SO_3 + HS = 2(FeO,SO_3) + HO,SO_3 + S.$$

Potash, soda, and ammonia throw down the hydrate of the sesquioxide, insoluble in excess; the presence of fixed organic acids prevents the precipitation altogether.

The carbonates of the alkaline earths throw down completely the salts of the sesquioxide.

Phosphate of soda precipitates the white phosphate of iron, which dissolves in ammonia.

The phosphate of the sesquioxide of iron is insoluble in acetic acid.

Ferrocyanide of potass, blue precipitate, ferrocyanide of iron, insoluble in HCl.

Ferricyanide, no precipitate, but a darker colour.

The sulphocyanide of potash gives a deep red colour to the most dilute solutions of the sesquioxide of iron.

The most dilute solutions of the peroxide give a bluish-black colour with an infusion of gall nuts.

NICKEL.

Symbol, Ni; *Equivalent*, 29·6.

This metal is found in meteoric stones, and may be prepared from the oxalate. The oxalate is obtained by precipitating the sulphate with oxalic acid. It is white, malleable, and strongly magnetic. It is prepared like cobalt.

The *Protoxide of Nickel*, NiO, has basic properties. It may be prepared by precipitating any soluble salt of the metal with caustic potash, and washing, drying, and igniting the precipitate, which is of a green colour. This oxide completely neutralises acids. The hydrated salts have a green colour: the anhydrous salts are yellow.

The *Sesquioxide of Nickel*, Ni_2O_3, is a neutral compound. It is produced when any salt of nickel is mixed with a solution of bleaching powder. This oxide is decomposed by heat, and evolves chlorine when heated with hydrochloric acid.

The *Chloride of Nickel*, $NiCl$, is prepared by dissolving the carbonate of nickel with hydrochloric acid, $NiO,CO_2 + HCl = NiCl + HO,CO_2$. From the green solution crystals of the chloride of nickel are obtained. When the crystals are anhydrous, they have a yellow colour.

Sulphate of Nickel, $NiO,SO_3 + 7HO$.—This is the most important of the nickel salts. The crystals are of a green colour, and usually contain seven equivalents of water. With potash and ammonia it forms double salts: $NiO,SO_3 + KO,SO_3 + 7HO$; $NiO,SO_3 + NH_4O,SO_3 + 6HO$.

Carbonate of Nickel, NiO,CO_2.—When a solution of the carbonate of soda is added to the chloride of nickel, a pale green precipitate of the carbonate and hydrate is obtained: $NaO,CO_2 + NiCl = NiO,CO_2 + NaCl$. The anhydrous salts of nickel are of a straw colour. When associated with water they are of a green colour. Nickel is chiefly employed in the manufacture of German silver. The best German silver contains 8 parts of copper, 3 of nickel, and 3½ of zinc; inferior kinds contain 8 of copper, 2 of nickel, and 4 of zinc.

Tests.

Hydrosulphuric acid and sulphide of ammonium have a similar reaction on nickel solutions to those described under cobalt.

Ammonia in neutral solutions precipitates a portion of hydrate of the protoxide, which is dissolved in excess, giving rise to a blue solution. If caustic potash be added to this solution, a green hydrate is thrown down.

The alkaline carbonates give a straw-coloured precipitate, easily soluble in carbonate of ammonia, the solution becoming of a bluish-green colour. Carbonates of the alkaline earths, same as cobalt.

Cyanide of potassium, a greenish-white precipitate, soluble in excess. From this solution the cyanide may be thrown down by hydrochloric or sulphuric acid, but the cyanide of cobalt cannot be again precipitated with acids.

In the outer flame of the blowpipe, compounds of nickel, when heated with borax and phosphorus salt, form a reddish-coloured globule; in the inner flame the metal is reduced, and the bead becomes of a grey colour.

COBALT.

Symbol, Co; *Equivalent*, 29·55.

Cobalt and *Nickel* are very similar in their general and chemical properties, and are often found in nature associated together. Cobalt is most conveniently obtained by heating the oxalate to a white heat, in a crucible lined with plumbago. It is a white brittle metal, and melts at a temperature of about 2000°. It does not oxidise in air at common temperatures. Dilute hydrochloric and sulphuric acids have but slight influence on the metal. It has strong magnetic properties.

The *Protoxide*, CoO, is a grey powder, very soluble in acids, and has strong basic properties, giving a fine red tint to its salts. It may be prepared by precipitating the chloride of cobalt with carbonate of soda, $CoCl + NaO,CO_2 = NaCl + CoO,CO_2$, and igniting the precipitate. When a solution of cobalt is mixed with caustic potash, a blue precipitate falls, which becomes violet on the application of heat; and then a dirty red colour from the absorption of oxygen, and its subsequent change into a hydrate.

The *Sesquioxide of Cobalt*, Co_2O_3, is a neutral oxide of a black colour, and insoluble. If the chloride of lime be mixed with a solution of protochloride of cobalt, the sesquioxide is obtained.

The *Chloride of Cobalt*, CoCl, is prepared by precipitating the protoxide with hydrochloric acid, $CoO + HCl = CoCl + HO$. It gives a deep red colour to the solution, which deposits, if concentrated, beautiful red crystals. When the solution is evaporated, it deposits blue anhydrous crystals, which again become red in contact with water. A dilute solution of this salt forms a sympathetic ink. Letters written with this ink are invisible, until the salt has been rendered anhydrous by exposure to heat, when the letters appear blue. When laid aside, moisture is absorbed, and the writing again becomes invisible.

Sulphate of Cobalt, $CoO,SO_3 + 7HO$.—This salt is similar in constitution with the sulphate of magnesia. It combines with the sulphate of potash and soda, forming double salts. If a solution of the sulphate of cobalt be mixed with oxalic acid, the oxalate of cobalt separates after some time.

Carbonate of Cobalt.—The alkaline carbonates precipitate a pale peach coloured carbonate and hydrate of cobalt, $3(CoO,HO) + 2(CoO,CO_2)$.

The oxide of cobalt gives a beautiful blue colour to glass. Smalt, or azure blue, a substance much used by painters, paper-stainers, and for giving a blue colour to pottery, is prepared in Saxony and Bohemia by a complicated process from the native arsenio-sulphide of cobalt.

Tests.

All the cobalt salts, when anhydrous, are blue; when in dilute solutions of water, they are of a crimson colour. Hydrosulphuric acid, in the presence of an alkaline acetate and the sulphide of ammonium, precipitates the black sulphide of cobalt, insoluble in excess of the precipitant, the sulphide of ammonium. In neutral solutions, free from the salts of ammonia, ammonia gives a blue precipitate, which after some time, changes to red, then green, then brown. If the solution contain any salts of ammonia, the solution becomes red, absorbs oxygen from the air, and the brown hydrate of cobalt is precipitated. Cobalt solutions, free from common air, give a blue precipitate with the fixed caustic alkalies, which on boiling becomes of a rose colour.

The alkaline carbonates precipitate a pink carbonate of cobalt.

The carbonates of the alkaline earths give no precipitate unless heated. Cyanide of potassium gives, with acid, solutions of cobalt, a brown white precipitate of the cyanide of cobalt, soluble in excess. Borax and phosphorus salts, when heated with any cobalt salt, give a blue glass globule in both blowpipe flames.

CHROMIUM.

Symbol, Cr; *Equivalent*, 28.

This metal is found as an oxide in combination with the oxide of iron. It is rather abundant in the Shetland Isles. Chromium was first obtained in a state of purity from the chromate of lead. The metal is obtained by mixing the oxide of chromium with $\frac{1}{6}$ its weight of charcoal powder, and placing the mixture in a crucible lined with charcoal. The crucible is then placed in a furnace, and raised to the very highest heat. The oxide of chromium is reduced by the charcoal, and the metal is obtained as a greyish-white brittle substance, and very difficult of fusion. The most powerful acids have little effect on chromium, and the pure metal has been seen by very few persons.

There are seven oxides of chromium, resembling very much the corresponding oxides of iron.

Protoxide of Chromium, CrO.—If potash be added to a solution of the protochloride of chromium, a brown precipitate is obtained, which passes into a deep red colour. If ammonia be substituted for potash, a pale green precipitate is obtained, which is the protoxide of chromium, CrO. This oxide is a powerful base, forming pale blue salts, which have a great affinity for oxygen.

Protosesquioxide, CrO,Cr_2O_3.—This compound is produced by the action of water upon the protoxide, with the evolution of hydrogen. This compound corresponds to the magnetic oxide of iron.

If the nitrate of the suboxide of mercury be mixed with the chromate of potash, and raised to red heat in a crucible, the chromate of mercury is formed, and the sesquioxide of chromium remains behind. It is used in painting and glass-staining. It has a fine green colour. After ignition, it is insoluble in acids. The hydrate of the sesquioxide, $Cr_2O_3,3HO$, may be prepared by boiling a dilute solution of the bichromate of potash with excess of hydrochloric acid, to which a small quantity of alcohol is every two or three minutes added, till the liquid changes from orange to green. From this solution the hydrate is precipitated by ammonia as a pale green colour.

The anhydrous sesquioxide may be prepared by heating to redness the bichromate of potash in an earthen crucible. The sesquioxide of chromium is a feeble base, resembling the sesquioxide of aluminum. It communicates a beautiful green tint to glass, and is used in enamel painting.

Protochloride, $CrCl$, is obtained by heating to redness the sesquichloride in a porcelain tube through which a current of pure dry hydrogen is passing. HCl is liberated; a white mass is obtained which dissolves in water with a great increase of temperature, yielding a blue solution which absorbs oxygen by exposure, and passes into a deep green colour. The protochloride of chromium is one of the most powerful reducing or deoxidising agents known.

Sesquichloride, Cr_2Cl_3.—This compound may be obtained anhydrous, by heating to redness in a porcelain tube the sesquioxide of chromium, charcoal, and dry chlorine gas passed through the tube. The sesquichloride condenses in the cold part of the tube in plates of beautiful violet colour. This compound is quite insoluble in water.

Chromic acid, CrO_3.—When the sesquioxide of chromium is heated with potash in contact with air, oxygen is absorbed, and CrO_3 generated. If 100 measures of a saturated solution of the bichromate of potash be mixed with 150 measures of sulphuric acid, and allowed to cool, the chromic acid crystallises out in brilliant red prisms. Chromic acid is very soluble in water, and is instantly reduced by contact with organic matter.

The *Chromate of Potash*, KO,CrO_3, is the source, from which all the chromic preparations are obtained. It is made from chrome iron ore, which is a compound of iron, $FeO + Cr_2O_3$. It is reduced to powder, calcined and heated for a long time in a furnace with the nitrate of potash, KO,NO_5. The product is treated with water, which yields a yellow solution, which deposits, on evaporation, crystals of the same colour. If sulphuric acid be added to the chromate of potash, KO,CrO_3, one half of the base is removed, and the neutral chromate converted into the bichromate, $KO,2CrO_3$. This salt is extensively used in the arts and, by slow evaporation, crystallises in beautiful red crystals. It melts when heated, is soluble in ten parts of water, and has an acid reaction.

Chromate of Lead, Pb,O,CrO_3. — This is the chrome yellow of the painter. When a solution of the nitrate of lead is mixed with a solution of chromate of potash, a yellow precipitate falls, which is the chromate of lead, or the chrome yellow of the powder. If this compound be boiled with lime-water, half of the acid is removed, and a sub-chromate formed, which has an orange-red colour. The colour of this sub-chromate rivals vermilion in the beautiful delicacy of the tint. The yellow and orange chrome colours are fixed on cloth by the alternate application of the two solutions.

Chromate of Silver, AgO,CrO_3.—When the nitrate of silver is mixed with the chromate of potash, the chromate of silver is precipitated as a reddish-brown powder.

Tests.

Lead salts give a lemon-yellow precipitate of the chromate of lead.

Silver salts, purple red precipitate.

The chromium compounds impart a beautiful green colour to borax in the inner and outer flame of the blowpipe.

Salts of the sesquioxide of chromium may be easily recognised:—

The caustic alkalies precipitate the hydrated oxide, easily soluble in excess of the reagent.

Ammonia precipitates a bluish-green hydrate, which is slightly soluble in strong ammonia, forming a pink solution.

The alkaline carbonates throw down a green precipitate.

Sulphuretted hydrogen, no change.

Sulphide of ammonia precipitates the hydrate of the sesquioxide.

A salt of chromium may be recognised by baryta salts, producing pale yellow precipitates.

ZINC.

Symbol, Zn; *Equivalent*, 32·55.

This metal is generally found in nature associated with lead ores, as the carbonate ZnO,CO_2, and sulphide ZnS. Our chief supplies of this metal are from Silesia. The ore is first roasted and then heated with charcoal to expel the water and carbonic acid.

It is then distilled at a red heat in a large earthenware retort. The reduced metal rises in vapour. This vapour is condensed in a suitable apparatus. The metal generally contains small portions of arsenic.

The metal is of a bluish colour, and is much used instead of lead, in consequence of its cheapness. It is crystalline, and under ordinary circumstances brittle; but raised to a temperature of 270° it is malleable, and may be rolled out in sheets without fracture. By such treatment the metal retains its malleability. At a temperature of 400° it may be reduced to powder. At a red heat it boils and volatilises, and in the presence of air burns with a green light, giving rise to the protoxide of the metal, ZnO.

Zinc is readily dissolved by dilute acids. It is generally used in the preparation of hydrogen.

The only known *oxide* of this metal is the protoxide ZnO. It is a strong base and may be prepared by raising the carbonate to a red heat.

The neutral carbonate of zinc, ZnO,CO_2, is found native; it is the chief ore of the metal and is called the calamine stone. It may be prepared by adding the carbonate of potash to a solution of the chloride of zinc, $ZnCl + KO,CO_2 = ZnO,CO_2 + KCl$.

The *Sulphate of Zinc*, $ZnO,SO_3 + 7HO$, *White Vitriol*, resembles in appearance the sulphate of magnesia. This salt forms the chief part of the residue left when hydrogen is evolved by the action of dilute sulphuric acid on zinc. The sulphate of zinc is used in medicine as an emetic. It forms double salts with potash and ammonia, $ZnO,SO_3 + KO,SO_3 + 7HO$.

The *Chloride of Zinc*, $ZnCl$, is prepared by dissolving the oxide of zinc in hydrochloric acid, $ZnO + HCl = ZnCl + HO$. It is very fusible and soluble in water and alcohol. The chloride of zinc combines with the chloride of ammonium to form a salt used in tinning and soft soldering the metals. The chloride of zinc is used as a disinfectant, and wood saturated with a solution of the chloride is not affected with dry rot. The salts of zinc are all soluble in hydrochloric acid, and most of them are soluble in water. In acid solutions hydrosulphuric acid produces no precipitate, but if an alkaline acetate be added, the white sulphide of zinc is thrown down. This precipitate is insoluble in potash.

Tests.

Sulphide of ammonium, precipitates the sulphide of zinc. Caustic alkalies give a white precipitate of the hydrate, soluble in excess.

The carbonates of the fixed alkalies precipitate the basic carbonate, insoluble in excess, but soluble in any salt of ammonia.

Carbonate of ammonia precipitates the carbonate, soluble in excess.

The carbonates of the alkaline earths; give no precipitate with zinc salts.

Zinc compounds when fused with carbonate of soda on charcoal in the inner flame of a blowpipe, give a white incrustation of the oxide of zinc. When moistened with cobalt and strongly heated, the colour becomes green.

CADMIUM.

Symbol, Cd ; *Equivalent*, 56.

This metal is more volatile than zinc. It is generally associated with its ores. When the calamine stone is subjected to distillation with charcoal the cadmium first comes off in vapour. It has a white tin colour, very malleable, burns when strongly heated. Hydrochloric and sulphuric acid have no effect on the metal at ordinary temperatures. It dissolves easily in nitric acid, and slightly oxidises in air.

The *Protoxide of Cadmium*, CdO, may be prepared by igniting the carbonate in a crucible; it is infusible, dissolves in acids, and its salts are without colour.

The *Sulphate of Cadmium*, $CdO,SO_3 + 4HO$.—This salt is obtained by dissolving the oxide in dilute sulphuric acid, $CdO + HO,SO_3 = HO,CdO,SO_3$. It is very soluble in water and forms double salts with ammonia and potash.

The *Sulphide of Cadmium*, CdS, is obtained by passing sulphuretted hydrogen gas through a solution of the chloride of cadmium, $CdCl + HS = CdS + ClH$.

Tests.

Hydrosulphuric acid and the sulphide of ammonium precipitate the yellow sulphide of cadmium: insoluble in alkalies.

Ammonia precipitates a white hydrate, soluble in excess, from which the hydrosulphuric acid precipitates the sulphide.

The fixed caustic alkalies produce a precipitate of the white hydrate, insoluble in excess.

The alkaline carbonates precipitate the carbonate of cadmium: insoluble in excess.

Cadmium compounds, when exposed to the inner blowpipe flame with carbonate of soda or charcoal, give a yellow-brown incrustation of the oxide of cadmium.

TIN.

Symbol, Sn; *Equivalent*, 58.

Tin occurs in nature as the oxide and sulphide, but chiefly as the former. The chief sources of this metal are the mines of Cornwall, Saxony, Bohemia, and Malacca. It is also found as an alluvial deposit in the beds of small rivers. These two varieties are called mine tin and stream tin. The latter has a much higher commercial value.

Tin ore is reduced by first grinding it to powder, then roasting to get rid of the water, sulphur, and arsenic. It is then heated with coal, and the reduced metal is cast into large square blocks, a specimen of the metal is assayed, and the stamp of the Duchy of Cornwall placed on each block. Tin, when quite pure, is of a white colour, soft and malleable. It is not much affected at ordinary temperatures by exposure to water or air: when raised to its melting point it rapidly oxidises, passing into a white powder, a substance which is used in the arts for polishing, and known as putty powder. The metal is easily dissolved by hydrochloric and nitric acid.

The *Protoxide of Tin*, SnO, is prepared by adding a solution of carbonate of potash to a solution of the protochloride of tin:—

$$SnCl,KO,CO_2 + HO = KCl + HO + CO_2 + SnO.$$

When the precipitated hydrate is dried and heated in an atmosphere of carbonic acid, it becomes a heavy black powder, unalterable in air, but burns like tinder in contact with a hot body, and changes into the binoxide. The protoxide has no basic properties.

Sn_2O_3.—If a solution of the protochloride of tin be mixed with a solution of the hydrated sesquioxide of iron, a grey substance falls, which is the sesquioxide of tin.

The *Binoxide of Tin*, SnO_2, has different properties according to the method by which it is prepared. If nitric acid be made to act on pure metallic tin, a white hydrate of the binoxide of tin is produced, insoluble in acids. If the bichloride of tin be precipitated by an alkali, a white precipitate is obtained of the hydrate of the binoxide, freely soluble in acids.

These precipitates have the same composition; and when heated lose their water and leave the pure binoxide of tin. They dissolve in caustic potash, and are precipitated unchanged by acids. The binoxide is a feeble base.

The *Protochloride of Tin*, $SnCl$, is obtained by dissolving metallic tin in hot hydrochloric acid. It is soluble in water and used as a deoxidising agent for the salts of mercury and silver.

The *Bichloride of Tin*, $SnCl_2$, may be prepared by exposing metallic tin to the action of chlorine, or distilling 1 part of powdered tin with 5 parts of the protochloride of mercury. A solution of the bichloride of tin is much used as a mordant.

Protosulphide of Tin, SnS, is prepared by fusing metallic tin with an excess of sulphur, and strongly heating the residue. With hydrochloric acid gently heated it evolves hydrosulphuric acid.

Sesquisulphide of Tin, Sn_2S_3, may be prepared by gently heating the protosulphide with a third part of its weight of sulphur.

The *Bisulphide of Tin*, SnS_2, *Mosaic Gold*, is prepared by raising to a low red heat in a glass flask 12 parts of tin, 6 of the chloride of ammonium, and 7 of flowers of sulphur. Two of these substances sublime, the bisulphide of tin collecting at the bottom in brilliant gold-coloured scales, which are used as a substitute for gold.

If a solution of the terchloride of gold be added to a dilute solution of the protochloride of tin, a precipitate is obtained which is known as the purple of Cassius, which is used in enamel painting. Tinned vessels are extensively used for domestic purposes. Thin plates of iron are superficially covered with tin. Good pewter is tin hardened with antimony.

Salts of the protoxide of tin pass by exposure to the atmosphere into salts of the binoxide.

Salts of the protoxide are precipitated by hydrosulphuric acid and sulphide of ammonium as a deep brown protosulphide of tin soluble in excess of the sulphide of ammonium.

Ammonia and the carbonate throw down a white hydrate of the protoxide, insoluble in excess. The fixed caustic alkalies produce the same results.

There are two modifications of the binoxide of tin, the one insoluble and the other soluble. The hydrate of the binoxide, SnO_2,HO, is the insoluble compound; it is formed when nitric acid acts on metallic tin.

$SnO,NO_5+2HO=SnO_2,2HO+NO_4$.—This oxide combines with acids to form salts, which are insoluble in acids. When separated from the excess of acid, they dissolve in water, from which they may be again precipitated by the acid. After ignition the binoxide refuses to unite with acids.

The soluble modification of the binoxide is prepared by fusing the insoluble compound with an excess of caustic soda. It is easily soluble in acids.

Tests.

Protoxide of Tin.—The caustic alkalies precipitate the white hydrate, soluble in excess of the precipitant.

The carbonates of ammonia, potash, and soda produce a similar precipitate, insoluble in excess. Ammonia gives the same result.

Sulphuretted hydrogen and the sulphide of ammonia give a black precipitate of the protosulphide.

Binoxide of Tin.—The caustic alkalies precipitate the white hydrate, soluble in excess.

Ammonia and the alkaline carbonates produce a white precipitate of the hydrate, slightly soluble in excess. Sulphuretted hydrogen and sulphide of ammonia a yellow precipitate of the sulphide, soluble in excess.

ANTIMONY.

Symbol, Sb; *Equivalent*, 129·03.

This metal is found in combination with sulphur. The ore is heated with metallic iron or the carbonate of potash, which removes the sulphur. It is of a bluish colour, highly crystalline, having a metallic lustre. It is so brittle that it can be reduced to powder in a pestle and mortar. It melts below a red heat, and volatilises at a white heat. It is not acted upon by air at the common temperature. If strongly heated, it burns with a white flame, producing crystals of the teroxide of antimony.

There are three oxides of antimony.

The *teroxide of antimony*, SbO_3.—This oxide is feebly basic. It may be prepared by burning metallic antimony at the bottom of a large red-hot crucible. There is a sulphate, a nitrate, and an oxalate of this oxide. When the teroxide of antimony, SbO_3, is boiled with the bitartrate of potash, it is dissolved, and the solution, by careful evaporation, yields crystals of tartar emetic, $KO,SbO_3,C_8H_4O_{10} + 2HO$, the tartrate of antimony and potash.

There is an intermediate oxide of antimony, SbO_4, which is the ultimate product of oxidising the metal by heat, with exposure to the atmosphere. It is an indifferent body.

Antimonic Acid, SbO_5.—When strong nitric acid is made to act on metallic antimony, antimonic acid is produced. When heated carefully it becomes anhydrous, and of a pale straw colour. At a red heat it is decomposed. It is insoluble in water and acids. The acid so prepared is monobasic, MnO,SbO_5. If the pentachloride of antimony, $SbCl_5$, be decomposed with excess of water, hydrochloric acid and antimonic acid is produced. The acid so produced is bibasic, $2MO,SbO_5$ or KO,HO,SbO_5. This latter salt is the only precipitant for soda. It must be added with care, or no result is obtained.

The *Terchloride of Antimony*, $SbCl_3$, the *Butter of Antimony*.—This compound is obtained when hydrosulphuric acid is generated by the action of hydrochloric acid on the tersulphide of antimony. The residual solution is then distilled until each drop of the condensed vapour, on falling into the water of the receiver, produces a white precipitate. The receiver is then changed, and the process carried on. The air should be carefully excluded. Alkaline solutions decompose the terchloride, producing the teroxide. Powdered antimony takes fire in chlorine gas, forming $SbCl_3$.

The *Pentachloride of Antimony*, $SbCl_5$, corresponds nic acid. It is prepared by passing a stream of dry chlorine over metallic antimony, gently heated in a porcelain tube.

The *Tersulphide of Antimony*, SbS_3, is the native sulphide from which the metal is obtained.

The *Pentasulphide*, SbS_5, is unimportant.

There is also a compound of antimony and hydrogen, which is prepared by adding the teroxide of antimony and zinc to a solution of dilute sulphuric acid. Antimoniuretted hydrogen comes off, which burns with a green flame. When a cold plate is pressed into the flame, metallic antimony is deposited. Antimony is used in medicine; and, with lead, it forms the type-metal of the printer. It is too brittle to be used alone, but its alloys have numerous applications.

The salts of the teroxide of antimony are decomposed by the addition of a large quantity of water, a white basic salt being separated. Tartaric acid dissolves this compound, and, when present in the solution, prevents its formation.

Tests.

Hydrosulphuric acid precipitates from acid solutions of the teroxide the tersulphide of antimony, soluble in the sulphide of ammonium and potash, but insoluble in the bicarbonate of ammonia. Ammonia precipitates the teroxide of antimony, insoluble in excess.

The carbonate of ammonia produces the same result, with the evolution of carbonic acid.

The fixed caustic alkalies precipitate the teroxide, soluble in excess. The presence of tartaric acid interferes with the reaction.

Metallic zinc precipitates from salts of the teroxide black metallic antimony. In the presence of nitric acid the teroxide is separated.

Nitric acid precipitates antimonic acid from any of its compounds.

Hydrosulphuric acid throws down from acid solutions of antimonic acid an orange-yellow pentasulphide of antimony, SbS_5.

Compounds of antimony, exposed with carbonate of soda to the reducing flame of the blowpipe, yield globules of the metal, a portion of which volatilises at a high temperature, forming a white incrustation of the teroxide on the charcoal.

COPPER.

Symbol, Cu; *Equivalent*, 31·50.

This metal is sometimes found pure, but generally associated with sulphur and iron. Large quantities of this ore are obtained from the mines of Cornwall and South Australia. The ore is taken to South Wales to be reduced. The process of reduction is rather complicated. The ore is roasted in a reverberatory furnace. The sulphide of iron is converted at a high temperature into the oxide; but the sulphur still exists in combination with the copper. The roasted ore is then heated with silicious sand, which combines with the iron, forming a fusible slag which separates from the sulphide of copper. The iron, by repeating this process, is at last removed. The sulphide of copper then decomposes, giving up its sulphur and combining with the oxygen. The oxide is then reduced by charcoal.

Copper is a good conductor of electricity. It is of a reddish colour, and very malleable. In a dry atmosphere, it undergoes no change; in a moist atmosphere, it becomes covered with a green crust of the carbonate of the metal. Dilute hydrochloric and sulphuric acid act slowly upon copper. Boiling sulphuric acid and cold nitric acid dissolve it readily; the first yielding sulphurous acid, and the second, nitric oxide, or binoxide of nitrogen. There are two oxides of copper which form salts.

The *Protoxide*, CuO; and the *Suboxide*, Cu_2O.—The protoxide may be prepared by decomposing the nitrate in a crucible at a red heat. If the carbonate of copper in solution be mixed with a solution of caustic potash, a bulky blue precipitate of the hydrated oxide falls, which, on being boiled, becomes of a dark brown colour: this is the anhydrous oxide. The protoxide unites with all the acids, forming a series of important salts. The suboxide of copper (Cu_2O) often occurs, beautifully crystallised with copper ores. It may be obtained artificially by heating in a covered crucible 5 parts of the black oxide and 4 parts of copper filings, or by adding sugar to a solution of the sulphate of copper, and then adding caustic potash in excess. The sugar deprives the black protoxide of half its oxygen, the suboxide precipitating as a brown-red powder. The salts of suboxide are colourless and very unstable, owing to their tendency to absorb oxygen, and pass into higher oxides. The protoxide gives a green colour to glass: the suboxide, a red colour.

Sulphate of Copper, Blue Vitriol, $CuO, SO_3 + 5HO$.—The old copper sheathing of ships is calcined and dissolved in dilute sulphuric acid, or by oxidising the sulphide. This salt crystallises in large blue crystals, soluble in water; and, when anhydrous, is nearly of a white colour. At a high temperature the salt is decomposed. It combines with potash and ammonia, forming double salts of a blue colour.

Nitrate of Copper, $CuO, NO_5 + 3HO$.—This salt is prepared by dissolving the metal in nitric acid. It forms deep blue crystals very soluble in water.

Carbonate of Copper, CuO, CO_2.—Carbonic acid, in various proportions, combines with the oxide of copper. When a solution of the carbonate of soda is added in excess to a solution of the sulphate of copper, a pale blue precipitate is obtained. On heating, it becomes sandy, and of a green colour. It has the following composition: $CuO,CO_2 + CuO,HO + HO$. This salt is used in painting. The green mineral malachite has a similar composition. A natural crystalline compound of the carbonate of copper and water, having a deep blue colour which has not yet been imitated artificially, has the following composition: $2(CuO,CO_2) + CuO,HO$. Verditer is said to be similar in its composition.

Chloride of Copper, $CuCl + 2HO$, is easily prepared by adding hydrochloric acid to the protoxide of copper. It forms green crystals, very soluble in water and alcohol. When slightly heated, it loses its water of crystallisation, and becomes of a yellowish-brown colour. At a high temperature, it loses half its chlorine, becoming the subchloride, Cu_2Cl. This compound may also be formed by introducing copper filings and gold leaf into a jar of chlorine gas.

Arsenite of Copper (*Scheele's green*) is prepared by mixing a solution of the sulphate of copper with a solution of the arsenite of potash. It is a bright green powder, and is used in painting.

Brass is an alloy of 70 parts of copper and 30 parts of zinc. Gun-metal is an alloy of 90 parts of copper and 10 of tin. Bell-metal contains a larger proportion of tin. Good bronze consists of 91 parts of copper, 2 parts of tin, 6 of zinc, and 1 of lead. The brass of [illegible] contained tin instead of zinc.

Tests.

Hydrosulphuric acid, and the sulphide of ammonium, precipitate the black sulphide of copper.

The fixed caustic alkalies throw down the blue hydrate of copper, which, on boiling, loses its water and becomes black.

Ammonia produces a greenish blue precipitate, which dissolves in excess, producing a beautiful blue azure colour.

The carbonates of the fixed alkalies throw down a greenish-blue carbonate of copper, which, by boiling, becomes black in an excess of the precipitant.

Cyanide of potassium gives a yellow-green precipitate, soluble in excess.

Ferrocyanide of potassium produces a dark, purple red precipitate.

Copper may be thrown down in very dilute neutral solutions in a pure metallic state, by introducing a bright piece of metallic iron which becomes at once covered with copper. When heated upon charcoal with the carbonate of soda in the inner blowpipe flame, metallic copper is obtained.

LEAD.

Symbol, Pb; *Equivalent*, 103·56.

The most abundant ore of lead is the native sulphide (galena), which not unfrequently contains small quantities of silver. The ore is placed on the floor of a reverberatory furnace. A considerable portion of the sulphide becomes by oxidation converted into the sulphate. The ore is then thoroughly mixed together, the temperature raised, the sulphide and sulphate react upon each other, producing sulphurous acid and metallic lead:—

$$PbO, SO_3 + PbS = 2SO_2 + 2Pb.$$

Lead is very malleable, it has no elasticity, and very little strength. It melts at 600°, and volatilises at a white heat. In moist air it becomes coated with a grey film of the suboxide. When melted it absorbs oxygen, passing rapidly into the protoxide. All the acids, except nitric, act slowly upon lead. There are four oxides of lead.

Protoxide, Litharge, Massicot, PbO.—This oxide is obtained when melted lead is exposed to a current of hot air, or by heating the carbonate to a dull redness. When carefully prepared, it is of a pale yellow colour; when raised to a red heat, it fuses and forms crystals on cooling of a brick-red colour. It is slightly soluble in water, giving an alkaline reaction. In a melted state, it dissolves in silicious matter with great energy. It is largely used in the manufacture of glass, upon which it confers great brilliancy. It unites with acids, forming a large class of colourless salts, if the acid be without colour.

Red Oxide of Lead, Red Lead, $Pb_3O_4 = 2PbO + PbO_2$.—This oxide is prepared by exposing for a long time to a dull red heat the protoxide of lead which has not been fused. It is a brilliant red powder. By the action of acids it is converted into the protoxide and binoxide. It is used in painting, and sold as a substitute for vermilion.

The *Binoxide* or *Brown Oxide*, PbO_2, is obtained by digesting red lead in dilute nitric acid. At a red heat it is converted into the protoxide. Hydrochloric acid converts it into the chloride. When heated with sulphuric acid, it forms the sulphate of lead, one equivalent of oxygen being liberated. The binoxide of lead is used to separate sulphurous acid from gaseous mixtures, by converting the sulphurous acid into sulphuric,—$PbO_2 + SO_2 = PbO,SO_3$.

The *Suboxide of Lead*, Pb_2O, is obtained when the oxalate of lead is heated to a dull red heat in a retort. It absorbs oxygen with avidity, passing into the protoxide.

The *Nitrate of Lead*, PbO,NO_5.—If the carbonate of lead be dissolved in dilute nitric acid, the nitrate of lead is thrown down, and the carbonic acid liberated; or the protoxide may be dissolved in dilute nitric acid, and concentrated till the crystals separate.

Carbonate of Lead, White Lead, PbO, CO_2.—This salt is sometimes found crystallised in long white needles with other metallic ores. It is prepared by adding in the cold the carbonate of soda to the acetate of lead. It is, however, largely manufactured for painting in the following manner:—Into a great number of earthen jars, arranged side by side, a little vinegar is poured. These jars stand on manure or tan; to each jar, a roll of sheet lead is adapted, or bars cast into the shape of a portcullis are placed, so as not to touch the vinegar or project above the jar. The jars are closely covered with boards, and a layer of tan or manure is placed upon the top. Another row of pots is then placed on the top of these, covered in like manner. By this arrangement a pile of many rows is constructed. After the lapse of some days, the sheets or bars of lead are taken out of the pots, when it is found that they are in great part converted into the carbonate, which is made ready for use by washing and grinding. The nature of the change is not well understood: it is thought that an acetate of lead is first formed, which is decomposed by the carbonic acid of the decomposing manure.

If clean metallic lead be placed in pure water with exposure to the air, a white powder makes its appearance, covering the whole of the lead. This substance is slightly soluble, and may be easily detected. This deposit is chiefly due to the action of the carbonic acid dissolved in the water, and consists of a combination of the carbonate and hydrate. In common river and spring water, a little sulphate is almost always present; this forms a thin film of the insoluble sulphate of lead, which prevents further action. It is for this reason that leaden cisterns may be used without danger. If the water were quite pure, it would be contaminated with lead, and in course of time the whole of the lead would be dissolved. Water containing a large quantity of free carbonic acid cannot be safely kept in leaden cisterns or passed through leaden pipes. The carbonate, although very insoluble in pure water, is slightly soluble in water containing free carbonic acid.

The *Chloride of Lead*, $PbCl$.—This salt may be prepared by adding litharge to boiling dilute hydrochloric acid, filtering, and setting aside the filtrate to cool, or by mixing a saturated solution of the acetate of lead with the chloride of sodium.

Iodide of Lead, PbI.—This salt is obtained as a beautiful yellow precipitate when any soluble salt of lead is mixed with a solution of the iodide of potassium.

Tests.

Lead salts are thrown down as a black sulphide on the addition of hydrosulphuric acid or sulphide of ammonium. The insoluble compounds of lead, by digestion with sulphide of ammonium, yield the sulphide if hydrochloric acid is present. The precipitate approaches a brownish-red colour.

The soluble chlorides, as hydrochloric acid, precipitate in solutions of the oxide. The white chloride is soluble in an excess of water or potash.

Sulphuric acid precipitates the white sulphate.

Ammonia precipitates all the lead salts except the acetate, as the white hydrate, insoluble in excess.

The chromates precipitate the yellow chromate of lead from solutions of the oxide, insoluble in dilute nitric acid.

The fixed caustic alkalies precipitate the white hydrate, soluble in excess.

The alkaline carbonates precipitate the white carbonate, insoluble in excess, but soluble in caustic potash.

The compounds of lead, when heated in the blowpipe on charcoal with the carbonate of soda, yield a globule of metallic lead.

BISMUTH.

Symbol, Bi; *Equivalent,* 213.

This metal is of a reddish colour. It is separated from the earthy matter with which it is associated by heating the mixture in a crucible. It volatilises at high temperatures, is slightly oxidised by contact with air, and dissolves freely in dilute nitric acid.

The *Teroxide of Bismuth,* BiO_3, is the base of all the bismuth salts. It fuses at a high temperature, and forms a powerful flux with silicious substances.

Nitrate of Bismuth, $BiO_3,3NO_5 + 10HO$. — This salt is prepared by dissolving a saturated solution of bismuth in nitric acid. If the solution be left for some time, large crystals of the neutral nitrate of bismuth are deposited. There is a basic nitrate of bismuth, $BiO_3NO_5 + 2HO$. This salt was extensively used as a cosmetic, but it injures the skin. It is occasionally used in medicine. Eight parts of bismuth, 5 of lead, and 3 of tin form a fusible metal used for taking impressions from seals. Bismuth also enters into the composition of the best kinds of type.

Tests.

Hydrosulphuric acid and sulphide of ammonium precipitate a dark tersulphide of bismuth, BiS_3.

The alkalies precipitate a white hydrate, insoluble in excess.

The alkaline carbonates precipitate a white carbonate, slightly soluble in an excess of the precipitant.

Chromate potash precipitates a yellow chromate of bismuth, insoluble in caustic potash, but soluble in nitric acid.

Bismuth salts give, when heated upon charcoal with carbonate of soda in the inner flame of the blowpipe, a yellow incrustation.

MERCURY.

Symbol, Hg; *Equivalent*, 100.

This important metal has been known from a very early period. It is found native, but chiefly in combination with sulphur. This ore is sometimes called cinnabar. It is found chiefly in Spain, and large quantities, in a state of great purity, have recently been imported from California. The mercury is separated from the sulphur by heating the ore in an iron retort, with scraps of iron or lime, at a high temperature. The mercury rises in vapour, which passes over into a large chamber, where the mercury is condensed, and the sulphurous acid escapes. Mercury is imported in iron bottles, each holding seventy-five pounds. The iron bottles are sometimes used for preparing oxygen from the binoxide of manganese. Mercury is sometimes adulterated with tin and lead. Mercury, at ordinary temperatures, is a liquid. At 662° it boils, and passes into a colourless vapour of great density; at a temperature of —40° F. it solidifies. At ordinary temperatures, mercury undergoes no change. When near its boiling point it absorbs oxygen, passing into a dark red powder, which is the red protoxide, and highest oxide of the metal. At a temperature approaching red heat it is decomposed into metallic mercury and oxygen. It was by this method that Priestley discovered the presence of oxygen gas. Dilute hydrochloric and sulphuric acid have no effect on the metal at ordinary temperatures. Concentrated boiling sulphuric acid converts mercury into the sulphate of the protoxide of the metal, with evolution of sulphurous acid. Nitric acid, at the common temperature, dissolves mercury without difficulty, with evolution of the binoxide of nitrogen.

There are two oxygen compounds of mercury, the grey or suboxide, Hg_2O, and the red or protoxide, HgO. The salts of the latter are the most permanent in their character. The isomorphous relationship of mercury to the other metals is not well established.

The *Suboxide of Mercury*, Hg_2, is prepared by digesting calomel, Hg_2Cl, with a solution of caustic potash, or by adding caustic potash to the nitrate of the suboxide. It is a dark grey powder, insoluble in water. Mercury and chalk, and grey ointment are preparations of this oxide.

The *Protoxide of Mercury*, HgO, may be prepared by heating the nitrate of the protoxide until the acid is expelled, or by adding an excess of caustic potash to a solution of corrosive sublimate, $HgCl$. A bright yellow precipitate of the protoxide is thrown down. It is slightly soluble in water, and very poisonous.

Sulphate of the Suboxide of Mercury, Hg_2O,SO_3, is prepared by adding sulphuric acid to a solution of the nitrate of the suboxide.

Sulphate of the Protoxide of Mercury, HgO,SO_3, may be obtained by boiling equal weights of sulphuric acid and mercury, until the metal is converted into a white crystalline powder, which is the sulphate of the protoxide of the metal. The solution is evaporated to dryness. Water decomposes this salt, forming an insoluble yellow compound, called turbith mineral, $3HgO,SO_3$. By continual washing with hot water, the acid is entirely removed, leaving the original protoxide of the metal.

The *Nitrate of the Suboxide of Mercury*, $Hg_2O,NO_5 + 2HO$, is obtained by dissolving mercury in an excess of dilute nitric acid. If the mercury be in excess, a salt is deposited, containing $3Hg_2O,2NO_5 + 3HO$, which is decomposed by water. These salts, having a similar composition, may be easily distinguished by rubbing them in a mortar with a little chloride of sodium.

The nitrate of the suboxide gives nitrate of soda and calomel. The basic salt gives a black compound of the nitrate of soda, calomel, and protoxide of mercury.

The *Nitrate of the Protoxide of Mercury*, HgO,NO_5, is prepared by dissolving the red oxide in excess of nitric acid and gently evaporating. When the solution, which has this composition, $HgO,NO_5 + 2HO$, has become thick, it is placed under a bell jar over strong sulphuric acid, when large crystals are deposited: $2(HgO,NO_5) + HO$.

The *Subchloride of Mercury, Calomel,* Hg_2Cl, is a very important compound. It may be prepared by pouring a solution of the nitrate of the suboxide into a large excess of dilute solution of common salt. Calomel is usually prepared by the following method. The dry sulphate of the red oxide of mercury is rubbed in a mortar with the same weight of mercury the salt contains, and a quantity of the chloride of sodium. When a uniform mixture has been produced, it is placed in a retort. It sublimes at a temperature below a red heat. The vapour is conveyed into a chamber, and condensed in a finely divided state:—

$$HgO,SO_3 + Hg + NaCl = Hg_2Cl + NaO,SO_3.$$

Calomel is a white heavy insoluble powder. It is not affected with cold nitric acid, but hot nitric acid dissolves it. It is decomposed by lime water, with production of the suboxide. If calomel should contain any chloride of mercury it is very dangerous. This chloride may be detected by boiling the calomel with water, filtering the solution, and adding caustic potash. Any corrosive sublimate is detected by a yellow precipitate.

The *Protochloride of Mercury,* HgCl, may be prepared by dissolving the red oxide in hot hydrochloric acid, when crystals of corrosive sublimate separate as the solution cools. It is generally prepared by rubbing a mixture of the dry sulphate of the protoxide of mercury, with an equal weight of dry common salt:—

$$HgO,SO_3 + NaCl = HgCl + NaO,SO_3.$$

Corrosive Sublimate is a white crystalline compound. It absorbs ammoniacal gas freely, is soluble in water, and forms a series of double salts;—

$$HgCl + NH_4Cl + HO.$$

Organic substances containing nitrogen unite with corrosive sublimate, forming insoluble compounds. It is on this account that albumen is a good antidote to poisoning by corrosive sublimate. Hyam's patent for preserving timber, cordage, &c. consisted in saturating the substances with a solution of corrosive sublimate.

Subiodide of Mercury, Hg_2I, is obtained when a solution of the iodide of potassium is added to a solution of the nitrate of the suboxide of mercury. The subiodide separates as an insoluble dirty yellow precipitate.

The *Protoiodide of Mercury*, HgI, is prepared by adding, in equivalent proportions, a solution of the iodide of potassium to the protochloride of mercury. The precipitate is first yellow, but changes into a brilliant permanent scarlet colour.

Subsulphide of Mercury, Hg_2S, is thrown down as a black precipitate when a solution of hydrosulphuric acid is mixed with a solution of the nitrate of the suboxide of the metal. It is decomposed by heat into sulphur and metallic mercury.

The *Protosulphide of Mercury*, *Cinnabar*, *Vermilion*, HgS.—When hydrosulphuric acid gas is passed gradually into a solution of corrosive sublimate, a white precipitate is obtained, a compound of the salt and sulphide. An excess of gas converts it into the black sulphide. When this black sulphide is sublimed, it undergoes no change of composition, but it becomes a dark red crystalline compound known as cinnabar and vermilion. When the protosulphide is heated in air, it yields sulphurous acid and metallic mercury. The protosulphide is not attacked by alkalies or strong mineral acids. It dissolves in aqua regia.

Mercury dissolves many of the metals, as gold, silver, lead, tin. A mixture of tin and mercury is used in silvering glass. The compounds are called amalgams.

All the basic salts of the suboxide are white; the neutral salts are yellow.

Salts of the suboxide are reduced by zinc, iron, copper, and the sulphate of the protoxide of iron.

Tests.

Hydrosulphuric acid precipitates the black subsulphide, Hg_2S.

The sulphide of ammonium produces the same result, insoluble in excess. It is decomposed into metallic mercury and sulphide of mercury on the addition of potash.

The soluble chlorides and hydrochloric acid precipitate the white subchloride, which is converted into the black suboxide on the addition of any caustic alkali.

The alkaline carbonates precipitate the carbonate of the suboxide.

Salts of the protoxide are precipitated as the black protosulphide, on the addition of hydrosulphuric acid and the sulphide of ammonium. This precipitate is soluble in caustic potash if the sulphide of ammonium be in excess. On the addition of a small quantity of these reagents the precipitate is white; but on a greater addition the black sulphide is produced, insoluble in concentrated boiling nitric acid, but easily dissolved in nitro-hydrochloric acid.

Ammonia, and the carbonate of ammonia, give a white precipitate.

All the compounds of mercury, when heated in a test-tube with carbonate of soda slightly moistened, yield globules of metallic mercury easily seen through a lens. Mercury, in neutral and slightly acid solutions, may be detected on the introduction of a piece of bright copper foil. After a few minutes, a grey metallic covering of mercury is produced.

SILVER.

Symbol, Ag; *Equivalent*, 108.

This metal is found in its pure state, but chiefly as the sulphide, bromide, and chloride, and in combination with lead ores. The chief sources of this metal are the mines of the Hartz mountains in Germany, Kongsberg in Norway, and the Andes in America. The ores are too poor to admit of the metal being extracted by smelting, although fuel may be abundant in the neighbourhood. It is therefore obtained by a complicated process of amalgamation with mercury. Pure silver may be obtained by dissolving the metal in pure nitric acid. If copper be present, the solution will be of a faint green tint, If gold be present, a black undissolved powder will remain. If the solution be now mixed with hydrochloric acid, a white insoluble precipitate of the chloride of silver will be obtained. This is washed, dried, and mixed with twice its weight of anhydrous carbonate of soda, and heated in an earthen crucible, gradually raised in temperature until it approaches a white heat. The oxygen and carbonic acid are evolved, and an alloy of metallic silver and sodium is formed. The silver button at the bottom of the crucible is easily separated from the sodium. Silver is of a white colour, of great lustre, malleable, ductile, and an excellent conductor of heat and electricity. It is unalterable in air and water. When strongly heated in air or oxygen, it absorbs many times its volume of that gas; but when the metal solidifies it parts with its oxygen, giving a rough arborescent appearance to the surface of the metal. The addition of a small portion of copper prevents the absorption of oxygen. When silver is heated with powdered glass, it forms a silicate of an orange colour. Hydrochloric acid has little effect on silver. Hot sulphuric acid converts it into a sulphate. Nitric acid dissolves it easily. The metal has a great affinity for sulphur. Sulphuretted hydrogen tarnishes the surface with a black sulphide. This is the reason silversmiths usually burn their gas outside the shop.

There are three oxides of silver:—

The *Suboxide*, Ag_2O, which is unimportant.

The *Protoxide of Silver*, AgO, is a powerful base, neutralising the most powerful acids, which form colourless salts. It may be prepared by precipitating the nitrate of silver with caustic potash. It is soluble in ammonia and slightly soluble in pure water. At a red heat, it is decomposed into metallic silver and oxygen.

The *Peroxide* is unimportant.

The *Nitrate of Silver*, *Lunar Caustic*, $AgON,O_5$, is prepared by dissolving metallic silver in nitric acid. The solution is concentrated and evaporated. When melted and poured into cylindrical moulds it forms the lunar caustic of the surgeon. When exposed to light, the salt blackens, especially if any organic matter be present. It stains the skin black, and is employed as a dye for the hair, and the manufacture of marking ink for linen.

Sulphate of Silver, AgO,SO_3.—This salt is prepared by precipitating a concentrated solution of the nitrate of silver with sulphate of potash.

The *Hyposulphate of Silver*, $AgO,S_2O_5 + HO$.—This salt is of a white colour, very soluble and permanent.

The *Hyposulphite of Silver*, $AgO,S_2O_2 + HO$, is a very unstable compound, insoluble, and of a white colour. It combines with any alkaline hyposulphite, forming soluble compounds of a sweet taste. The oxide and chloride of silver give rise to similar compounds with the alkaline hyposulphites.

The *Carbonate of Silver*, AgO,CO_2, is prepared by mixing a solution of the nitrate of silver with a solution of the carbonate of soda.

Chloride of Silver, $AgCl$.—This salt is produced when any silver compound is precipitated with a soluble chloride. It is of a white curdy appearance, quite insoluble in water and nitric acid. It is sometimes found native, and is called horn silver. The chloride of silver is decomposed by exposure to light, especially if any organic substances are present. It is reduced to its metallic state when put into water with metallic zinc or iron. It is soluble in ammonia and the cyanide of potassium.

The *Iodide of Silver*, AgI, is obtained when a solution of the iodide of potassium is added to the nitrate of silver. It is nearly insoluble in ammonia. In this respect it is an exception to the silver salts generally. The bromide resembles the chloride.

The *Sulphide of Silver*, AgS, is obtained by precipitating a solution of silver with hydrosulphuric acid. The salts of silver are without colour, and leave in most cases metallic silver after ignition. Silver compounds are decomposed by zinc, iron, copper, sulphate of the protoxide of iron, and chloride of tin, metallic silver being precipitated.

Tests.

Hydrosulphuric acid and sulphide of ammonia precipitate the black sulphide of silver.

Hydrochloric acid and the soluble chlorides throw down a white curdy precipitate of the chloride of silver. Soluble in ammonia, but insoluble in nitric acid.

The fixed caustic alkalies, give a brown precipitate. Neutral solutions of silver salts are rendered turbid by ammonia.

The alkaline carbonates throw down the white carbonate of silver, soluble in an excess of the carbonate of ammonia.

PLATINUM.

Symbol, Pt; *Equivalent*, 98·7.

This metal in a crude state, associated with palladium, rhodium, iridium, osmium, and iron, is found among the diluvial accumulations on the slope of the Ural mountains in Russia, Ceylon, and South America. It has never been seen in the rocks of any of the materials accompanying it. It is extracted from the alloy by a very difficult and complicated process.

Platinum is a white metal, very malleable, ductile, and infusible, scarcely yielding at times to the oxy-hydrogen blowpipe. It is the heaviest known substance. Neither air, moisture, nor ordinary acids have the slightest effect on platinum. It is dissolved in nitro-hydrochloric acid, and oxidised by the fused hydrate of potash. Spongy platinum and platinum black have the property of condensing gases into their pores, especially oxygen. It is this property which is thought to determine the union of hydrogen and oxygen when a piece of spongy platinum is held before a jet of hydrogen.

There are two oxides of platinum.

The *Protoxide*, PtO, of platinum may be obtained by digesting the protochloride with caustic potash: a black powder, the protoxide of platinum is produced, soluble in an excess of potash.

The *Binoxide of Platinum*, PtO_2, is prepared by boiling the bichloride with an excess of caustic soda, then adding acetic acid. A bulky brown powder is produced which becomes black and anhydrous when gently heated.

The *Bichloride of Platinum*, $PtCl_2$, is the most important compound. This compound is formed when platinum is dissolved in nitro-hydrochloric acid. On evaporating to dryness, a reddish-brown residue is left, very soluble in water and alcohol. The solution has an orange-yellow colour. It forms a variety of double salts with metallic chlorides. When a solution of the bichloride of platinum is added to any solution of potash, a bright yellow crystalline precipitate is produced of the bichloride of platinum and the chloride of potassium, $PtCl_2.KCl$. A few drops of hydrochloric acid promote the formation of the precipitate.

The *Chloride of Platinum*, $PtCl$, is produced when sulphurous acid is passed into a solution of the bichloride until it ceases to give a precipitate with chloride of ammonium. It is a greenish powder insoluble in water, but soluble in hydrochloric acid.

Tests.

A solution of the binoxide or bichloride of platinum can be recognised by a yellow precipitate with chloride of ammonia and chloride of potassium. Hydrosulphuric acid, on the application of a moderate heat, precipitates the bisulphide, of a brownish-black colour, insoluble in nitric acid and hydrochloric acid, but soluble in the sulphide of ammonium and nitrohydrochloric acid. Chloride of tin, in the presence of free hydrochloric acid, gives a reddish-brown colour to solutions of platinum.

GOLD.

Symbol, Au; *Equivalent*, 98·33.

This metal is widely diffused throughout nature. Small quantities are found in iron pyrites and the granitic rocks. It is always found in a pure metallic state, beautifully crystallised in a cubic form, associated with sand, quartz, and oxide of iron. The sands of various rivers contain this metal in a finely-divided state. This is the gold-dust of commerce.

Gold does not oxidise in air or water. It is not affected by hydrosulphuric acid or any single acid. A mixture of hydrochloric and nitric acid is used to dissolve it. Gold melts at a temperature of 5237° Fahr. It is the most malleable and ductile of all the metals. The thinnest gold-leaf is said not to exceed the $\frac{1}{200000}$th of an inch in thickness.

There are two oxides and two chlorides of gold.

The *Protoxide of Gold*, AuO, is obtained by adding a solution of caustic potash to the protochloride. It is a green powder, slightly soluble. The solution is very unstable, and decomposes into metallic gold and the teroxide of the metal; the latter dissolves, and the former falls in fine powder to the bottom.

The *Teroxide of Gold*, AuO_3, is obtained by adding magnesia to the terchloride. It is a reddish-yellow powder, which dries into a dark brown colour.

The *Protochloride of Gold*, AuCl, is produced by evaporating the terchloride to dryness and exposing to a heat of about 440°, until chlorine ceases to be given off. It is a yellowish-white mass insoluble in water.

The *Terchloride of Gold*, $AuCl_3$, is the most important compound. It may be produced on a small scale by sending chlorine through water in which pieces of gold leaf are suspended. The gold dissolves, and if the solution be evaporated it yields crystals of the terchloride. It is always produced when gold is dissolved in nitro-hydrochloric acid. It is very deliquescent and soluble in water. The terchloride combines with other metallic chlorides forming double salts. The general formula of these compounds may be thus represented, $MCl + AuCl_3$.

The terchloride of gold, with an excess of the bicarbonate of soda or potash, is used for gilding small ornamental copper articles. They are first cleaned with dilute nitric acid, then boiled for some time in the solution.

Gold coins are alloyed with a proportion of silver or copper to increase its hardness and durability. English standard gold contains $\frac{1}{12}$ of copper.

Gilding on copper is now performed by electrolysis. The solution employed for this purpose is obtained by dissolving the oxide or cyanide of gold in a solution of the iodide of potassium.

The most important reducing agents for gold are the sulphate of the protoxide of iron, oxalic acid, and chloride of tin. The sulphate of the protoxide of iron precipitates metallic gold as an impalpable brown powder, which may be collected on a filter. Dried and rubbed with a smooth body, it assumes the yellow colour and metallic lustre of the metal.

The chloride of tin, containing a little bichloride, on the addition of a drop of nitric acid, produces in very dilute solutions of gold a purple precipitate, known as the purple of Cassius.

ARSENIC.

Symbol, As; *Equivalent*, 75.

This metal is of a steel-grey colour. It is sometimes found native, but more frequently in combination with oxygen, sulphur, iron, nickel, and cobalt. These arsenides of the metals are placed in a reverberatory furnace, and the volatile products condensed in a chimney of peculiar construction. The crude arsenious acid is then purified by sublimation. It is then reduced with charcoal in an earthenware retort, and the metal easily sublimes.

Arsenic, when heated, oxidises into arsenious acid. It has the odour of garlic, and combines with the metals, like sulphur and phosphorus. It unites with oxygen in two proportions: — arsenious acid, AsO_3, and arsenic acid, AsO_5. There is no basic compound of this metal.

Arsenious Acid, White Oxide of Arsenic, AsO_3, is prepared by heating arsenic in contact with air. It is very soluble in hot water, and the acid crystallises in transparent octohedrons. The alkalies dissolve this compound readily, forming arsenites. Arsenious acid is a powerful poison. In experimenting with arsenic compounds, the student should avoid inhaling the vapours.

Arsenic Acid, AsO_5.—This compound is produced by dissolving arsenious acid in hot hydrochloric acid, and oxidised with nitric acid so long as red vapours are produced. The solution is then carefully evaporated to dryness. The acid so obtained is white and anhydrous. It slowly and entirely dissolves in water, the solution being slightly acid, which, on being evaporated, yields, after some time, crystals of hydrated arsenic acid. When heated, it is decomposed into arsenious acid and oxygen.

Arsenic acid resembles in the closest manner phosphoric acid, forming salts corresponding to the phosphates.

The *Arseniate of Soda*, $2NaO,HO,AsO_5 + 24HO$, cannot be distinguished in its appearance from the phosphate of soda. It may be prepared by adding a solution of carbonate of soda to a solution of arsenic acid until the acid is completely neutralised, and then evaporating.

There are also two other arseniates of soda, having the following composition:—$3NaO,AsO_5 + 24HO$, and $NaO,2HO + AsO_5,24HO$.

There are three sulphides of the metal.

Realgar, AsS_2, which is found native, and can be formed artificially by heating arsenic acid with the equivalent proportion of sulphur. It is employed in painting and for making white fire.

Orpiment, AsO_3, is also a natural compound. This may be prepared by precipitating an acid solution of arsenic with hydrosulphuric acid; after the lapse of some time a higher sulphur compound is precipitated.

The *Pentasulphide of Arsenic*, AsS_5, is not important.

The three sulphur compounds of arsenic are all sulphur acids.

Tests.

Arsenious acid can be detected by characters which cannot be mistaken.

The nitrate of silver, added to a solution of arsenious acid and water; on the addition of a drop or two of ammonia, a yellow precipitate of the arsenite of silver is thrown down, very soluble in ammonia and nitric acid.

Sulphate of copper, on the addition of a little alkali, added drop by drop, produces a yellow green precipitate (Scheele's green), soluble in ammonia.

Sulphuretted hydrogen, passed through a solution of arsenious acid, previously acidulated with a few drops of hydrochloric acid; a copious bright yellow precipitate of orpiment is thrown down, soluble in ammonia, and can be reprecipitated by sulphuric acid.

The detection of arsenious acid, when mixed with the contents of the stomach, is much more difficult. The presence of organic matter so interferes with the action of the tests just given, as to render their results worthless in a matter of so much importance.

Marsh's plan is, perhaps, the most simple and delicate. Slips of pure zinc are placed in the lower bulb, by unscrewing the stop-cock. The suspected liquid, which has been acidulated with pure sulphuric acid in the proportion of one ounce of acid to a pint of the liquid, is then introduced among the scraps of zinc. The stop-cock is closed, and hydrogen is generated in the lower bulb, which forces the liquid into the upper bulb. The apparatus now acts by hydrostatic pressure. As soon as the stop-cock is opened the gas is forced through the jet and inflamed; a piece of white porcelain is held in the flame, when any arsenic is at once recognised by the presence of a black, brilliant metallic spot of reduced arsenic on the porcelain.

The evolution of hydrogen in the lower bulb combines with the arsenious acid, and the gas which is inflamed is arseniuretted hydrogen, which in contact with the cold porcelain is reduced to a metallic state.

RÉSUMÉ.

POTASSIUM.

Symbol, K; *Equivalent*, 39.

Discovered by Sir H. Davy in 1807.

Physical Properties.—A white, soft metal, capable of being cut with a knife.

Specific Gravity.—0·86.

Chemical Properties.—A simple body absorbing oxygen at all temperatures, and decomposing water in the cold.

Preparation.—By reducing the carbonate of potash and charcoal to powder, and heating in an iron retort, the potassium in a melted state falls into a receiver containing naphtha.

PROTOXIDE OF POTASSA, POTASH.

Symbol, KO; *Equivalent*, 47.

Physical Properties.—A white, solid body. When touched with the fingers it feels soapy. It is fusible at a dull red heat, and volatilises at a high temperature. Potash unites with water, forming the hydrate of potash, which is the true caustic potash, KO,HO. It is a very powerful alkali, completely neutralising the strongest acids. On exposure to the air potash absorbs moisture, becomes slightly liquefied; and when converted into the carbonate it melts.

Preparation. — The potash of commerce is an impure carbonate, and is chiefly obtained from wood ashes.

NITRATE OF POTASH, NITRE.

Symbol, KO,NO_5; *Equivalent,* 101.

Physical Properties. — A white salt, crystallised in long six-sided prisms. It is decomposed at a high temperature.

Chemical Properties. — It is soluble in water, and becomes liquid in a moist atmosphere. It is a poison when taken in large quantities. With sulphur and carbon it forms gunpowder.

Nitre is found in many parts of the East as a white efflorescence on the surface of the earth. The French manufacture it artificially from urine, chalk, and decomposing animal substances.

Powder differs in its composition according to its use. The following may be taken as the general composition: —

$$KO,NO_5 + S + 3C.$$

When powder has to be manufactured for mining or sporting, these proportions vary.

SODIUM.

Symbol, Na; *Equivalent*, 23.

Discovered by Sir H. Davy in 1807.

Specific Gravity.—0·97.

Physical Properties.—A white, silver-looking metal, solid at ordinary temperatures, and volatilises at a red heat.

Chemical Properties.—Decomposes warm water, and absorbs oxygen at all temperatures. It is kept in naphtha.

Preparation.—By heating the carbonate of soda and charcoal in iron retorts, the same as described under potassium.

SODA.

Symbol, NaO; *Equivalent*, 31.

Soda is very analogous in all its properties to potash.

The soda of commerce is an impure carbonate, chiefly obtained from the ashes of marine plants.

SULPHATE OF SODA.

Symbol, NaO,SO_3; *Equivalent*, 71.

Physical Properties.—A white salt, crystallised in four sides; soluble in twice its weight of cold water. It is used in the manufacture of hydrochloric acid and the carbonate of soda.

CHLORIDE OF SODIUM, COMMON SALT.

Symbol, NaCl; *Equivalent*, 68·5.

Physical Properties.—A white salt, crystallised in the form of cubes. It is soluble in water, and is obtained by evaporation. Our chief supplies are from the salt mines of Cheshire.

The brine springs of Cheshire yield a large supply. It is sometimes obtained by evaporating sea-water.

AMMONIA.

Symbol, NH_4O; *Equivalent*, 26.

The compound metal ammonium, of which ammonia is the oxide, has never been isolated.

Chemical Properties.—Ammonia has a powerful odour, and very volatile. It is strongly alkaline, and can be made to replace potash or soda. It is this property of ammonia which gives rise to a series of salts analogous to those of soda and potash.

The salts of ammonia contain an equivalent of water. Hypothetical formula of ammonium, NH_4. The hydrate is the oxide NH_3HO or NH_4O. All the salts of ammonia are constituted on this theory.

CARBONATES OF AMMONIA.

The neutral anhydrous carbonate of ammonia, NH_3,CO_2, is prepared by the direct union of carbonic acid with ammoniacal gas. It is very soluble in water.

The strong smelling salts of ammonia are prepared by subliming a mixture of sal ammoniac with chalk. It is probably a sesquicarbonate, $2NH_4O,3CO_2$.

CHLORIDE OF AMMONIUM, SAL AMMONIAC.

Symbol, NH_4Cl; *Equivalent*, 53·5.

A white, fibrous, crystalline body; volatile at a red heat, and soluble in water and alcohol. It is prepared by acting on the carbonate of ammonia with hydrochloric acid:—

$$NH_4O,CO_2 + HCl = NH_4Cl + HO,CO_2.$$

BARIUM.

Symbol, Ba; *Equivalent*, 68·5.

A white, brilliant metal, but speedily oxidises on exposure to air. It is very heavy, poisonous, and slightly soluble in water.

BINOXIDE OF BARIUM.

Symbol, BaO_2; *Equivalent*, 84·5.

Is of a greenish white colour and amorphous. It can be obtained by heating to redness baryta, BaO, in a current of oxygen or air. The binoxide is deoxidised into baryta at red heat. This alternate oxidation and deoxidation of baryta has been proposed as an economic source of oxygen.

CALCIUM.

Symbol, Ca; *Equivalent*, 20.

A white metal, discovered by Sir H. Davy in 1809, similar to Barium. Calcium has been seen by very few persons. On exposure to air it passes into the protoxide of calcium, CaO (lime).

LIME.

Symbol, CaO; *Equivalent*, 28.

Lime is a white, amorphous body, usually called quicklime. It absorbs moisture with great avidity, passing into the hydrate, CaO,HO, and on exposure to the air it absorbs carbonic acid, becoming the carbonate of lime.

It is more soluble in cold water than hot, forming a solution known as lime-water.

Lime is made from chalk or marble, by strongly heating in a limekiln, which drives off the carbonic acid.

With alumina, silica, and iron it forms hydraulic cements, which have the property of hardening under water.

Building mortar is a mixture of sand, lime, and water. It hardens on exposure to air. Good mortar ought to contain three parts of road or river sand to one of lime.

CARBONATE OF LIME.

Symbol, CaO,CO_2; *Equivalent* 50.

There are several varieties of carbonate of lime:

The crystallised forms, such as stalactite, spathose, and arragonite; amorphous, such as chalk and marble.

The pure carbonate of lime is without colour; insoluble in pure water, but soluble in water containing carbonic acid. It is decomposed by heat into carbonic acid and lime. Carbonate of lime is the source of incrustations in steam-boilers and tea-kettles. By heating water containing this salt the carbonic acid is driven off, and the insoluble carbonate of lime falls to the bottom. Stalactites in caverns are formed by the gradual percolation of water containing the carbonate of lime.

SULPHATE OF LIME, PLASTER OF PARIS, GYPSUM.

Symbol, CaO,SO_3; *Equivalent*, 68.

This salt is found in nature both in a crystallised and amorphous state. The anhydrous sulphate is the plaster of Paris; it absorbs water, and becomes a hard compact mass. It requires a red heat to deprive it of this water. The sulphate of lime is soluble in 500 parts of cold water. In its crystallised state alabaster is a familiar example.

ALUMINUM.

Symbol, Al; *Equivalent*, 13·7.

A silver-white metal, discovered by Wöhler in 1828, does not readily oxidise. Aluminum is chiefly obtained from a mineral called chrysolite, by reduction with potassium or sodium. (Specific gravity, 2·6.)

ALUMINA.

Symbol, Al_2O_3; *Combining Equivalent*, 51·4.

Alumina is a very weak base, and its salts have frequently an acid reaction. The ruby and sapphire are crystallised varieties of alumina, coloured with some metallic oxide. Emery powder, used for polishing glass and metals, is a coarse kind of crystallised alumina, called corundum.

ALUMS.

All the alums are double sulphates. Common alum may be taken as the type, which is a double sulphate of alumina and potash, $Al_2O_3,3SO_3 + KO, SO_3+24HO$. Alumina is used in the arts for dyeing and preparing skins.

IRON.

Symbol, Fe ; *Equivalent*, 28.

Pure iron is of a greyish-white colour, fusible at a high temperature. It has great tenacity, very ductile. Specific gravity 7·84. On exposure to the atmosphere, iron passes into the oxide ; at a red heat it decomposes water.

Oxides of Iron.—The protoxide Fe,O, which is a black powder, and very oxidisable. The sesquioxide Fe_2O_3, which is of a reddish-brown colour, and is obtained by precipitating, with ammonia, any of the salts of the protoxide of iron. The magnetic oxide Fe_3O_4, which is found in nature, and is the product of burning iron wire in oxygen and ferric acid, FeO_7.

ZINC.

Symbol, Zn ; *Equivalent*, 32.

This metal is first mentioned by Paracelsus, who died in 1530. It is of a bluish-white colour, soft, slightly flexible, and at a temperature of 270° very malleable. At a red heat it boils and volatilises, and if air be present, burns with a green flame, generating the oxide of zinc. Zinc has a specific gravity of 7.

Zinc is much used for roofing, instead of lead. It is durable, but inferior in this respect to lead. Cisterns, pails, and many domestic utensils are now made of zinc.

The protoxide of zinc, ZnO, is a white body, and a powerful base. It may be obtained by the direct oxidation of zinc. It is used as a bad substitute for white lead in painting.

Sulphate of zinc, ZnO,SO_3+7HO, or white vitriol. This salt is isomorphous with the sulphate of magnesia. It is of a white colour, soluble in water, but insoluble in alcohol. It is employed in calico printing.

TIN.

Symbol, Sn; *Equivalent*, 58.

Tin is a white silver-looking metal, has great metallic lustre, and is a good reflector of light. It has a specific gravity of 7·29. There are two important oxides of tin, the protoxide SnO, and the binoxide SnO_2. The salts of the protoxide are precipitated with chloride of gold. The precipitate is of a purple colour, and is known as the purple of Cassius.

The salts of the binoxide of tin give a white precipitate with sulphuretted hydrogen.

The salts of tin are used in dyeing. Thin sheets of iron dipped in molten tin, form the tin plate of commerce. A mixture of tin and lead forms the solder of the plumber. Tin and sulphur form mosaic gold. Pewter is tin hardened with antimony. Stereotype metal is composed of 6 of antimony and 1 of lead. Britannia metal is a compound of variable quantities of copper, tin, bismuth, and antimony.

COPPER.

Symbol, Cu ; *Equivalent*, 31·7.

Copper is of a red colour, very ductile and tenacious. When breathed on, it creates a disagreeable odour. It is fusible at a bright red heat, and has a specific gravity of 8·96.

On exposure to a moist atmosphere, copper becomes covered with a film of the carbonate of copper, which is of a green colour, and very poisonous.

There are four oxides of copper—two only are important. The suboxide Cu_2O, and the protoxide CuO. The suboxide is of a red colour, soluble in ammonia. The protoxide is the ordinary base of all the copper salts.

The salts of copper give, with ammonia, first a green, then a blue colour. The sulphate of copper CuO_2SO_3+5HO is of a blue colour, and poisonous. It is used in medicine and painting.

Brass consists of about 30 parts of zinc and 70 of copper. Copper is sometimes externally converted into brass by exposure, when at a red heat, to the vapour of zinc. The gold wire of Lyons is manufactured from copper bars prepared in this manner. Gun metal consists of 90 parts of copper and 10 of tin. Bell metal contains a larger proportion of tin. Bronze consists of 91 parts of copper, 2 of tin, 6 of zinc, and 1 of lead. The brass of the ancients was copper and tin.

LEAD.

Symbol, Pb; *Equivalent*, 103.

Lead is a bluish-white metal, having great lustre; on exposure to the air it becomes dull, owing to its absorption of oxygen; this thin film protects the metal from further oxidation. It is very malleable; but has little ductility and tenacity. The specific gravity of lead is 11·44. Lead melts at 600°; at a white heat it boils and volatilises. The salts of lead are without colour, except the chromates; they have a sweet taste, and are poisonous. When clean metallic lead is placed in pure water it becomes covered with a thin film of the hydrated protoxide of lead, formed by the dissolved oxygen of the water. This compound is slightly soluble in water. This deposit, however, is more frequently the carbonate in union with the hydrate of lead, and is a very insoluble compound. River and spring water almost always contain a little sulphate, which causes a thin coating of the sulphate of lead, and protects the metal from farther action. If water were quite pure, the lead would in time be dissolved; or if the water contain a large quantity of free carbonic acid, the water cannot be used with safety. The carbonate of lead, although insoluble in pure water, is slightly soluble in water containing free carbonic acid.

Soluble lead salts give, with sulphuric and hydrochloric acid, a white precipitate; with hydrosulphuric acid, a black precipitate.

White Lead, PbO,CO_2.—White lead is manufactured by placing sheets of lead in earthernware jars containing acetic acid. These jars stand in rows upon manure or tan. The acetic acid converts the lead into an acetate, which is subsequently converted into the carbonate by the carbonic acid of the decomposing manure. The process occupies some days. White lead is extensively used in painting.

MERCURY.

Symbol, Hg; *Equivalent*, 100.

This metal, at ordinary temperatures, is a liquid; at a temperature of 40°, it congeals; at a temperature of 662°, it volatilises. Mercury has a specific gravity of 13·59.

There are two oxides of mercury, and corresponding sulphides and chlorides,—the suboxide, Hg_2O, and the protoxide, or red precipitate, HgO.

Calomel, Hg_2Cl. — This is a white crystalline body, insoluble in water. It is very useful in medicine. Calomel may be prepared by decomposing the sulphate of the protoxide of mercury in the presence of free mercury, with a solution of common salt:—

$$HgO,SO_3 + Hg + NaCl = NaO,SO_3 + Hg_2Cl.$$

Corrosive Sublimate, $HgCl$, is of a white colour. This salt is soluble in water, to which it communicates a sweet metallic taste. With albumen, it forms a white insoluble precipitate. It is a powerful poison. White of egg is the best antidote.

Corrosive sublimate is prepared by decomposing the sulphate of the protoxide of mercury with the chloride of sodium:—

$$HgO,SO_3 + NaCl = NaO,SO_3 + HgCl.$$

Mercury combines with great facility with gold, silver, zinc, tin, lead and bismuth, forming alloys called amalgams. Mercury is employed to extract gold and silver from their ores; and for gilding, plating, and looking glasses. It is also used for thermometers and barometers.

SILVER.

Symbol, Ag; *Equivalent*, 108.

This is a very white, brilliant-looking metal, having great lustre. It is ductile and malleable; but not so much as copper. It has a specific gravity of 10·5, and melts at a bright red heat. It does not readily oxidise in air, is insoluble in water, but soluble in nitric acid.

Soluble salts of silver are generally without colour; but on exposure to the light they blacken. Hydrochloric acid and the soluble chlorides give a white precipitate, with silver solutions, soluble in ammonia.

Nitrate of silver is the lunar caustic of the surgeon.

The silver used for coin and plate in this country consists of an alloy of silver and copper, fixed by law; 111 parts of silver and 9 of copper is standard silver, and bears the Hall mark of the Goldsmith's Company. In France, three different standards are employed; one for currency, one for plate, and one for ornaments.

PLATINUM.

Symbol, Pt; *Equivalent*, 98·7.

This metal was discovered in 1752, and called white gold. Specific gravity 21. It is of a greyish colour, very difficult to fuse, and does not oxidise on exposure to the air.

Platinum can be forged and welded like wrought iron. It is made into crucibles and stills for the chemist and gold refiner.

GOLD.

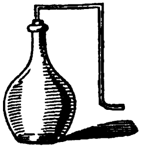
Symbol, Au; *Equivalent*, 197.

Gold is the most unalterable of all the metals. It is very ductile and malleable, the thinnest gold leaf not exceeding the $\frac{1}{200,000}$ of an inch in thickness. It dissolves in nitro-hydrochloric acid, producing the sesquichloride of gold, Au_2Cl_3, which is thrown down as a red-purple precipitate, with the protochloride of tin. Gold has a specific gravity of 19·5. Oxalic acid gives a brown precipitate.

Gold coinage in this country consists of an alloy in which 22 parts out of 24 are pure gold, and 2 parts of alloy. Good jewellery contains 18 parts out of the 24, and bears the mark of the Goldsmiths' Company. Jewellery with less than this proportion does not receive the mark. It not unfrequently happens that chains, rings, &c., sold as gold, do not contain more than 40 per cent. of the pure metal.

APPARATUS.

1.



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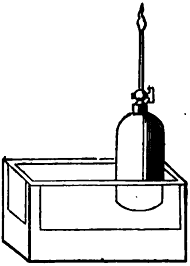

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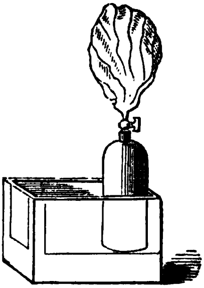

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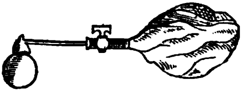


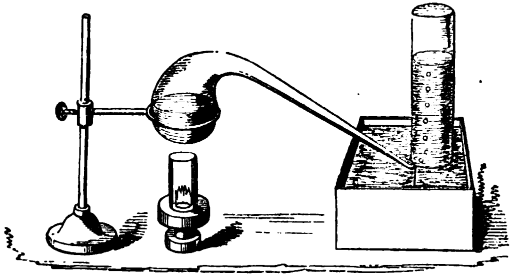
1. Arrangement for preparing hydrogen and gases not requiring heat. The delivery tube is placed in the pneumatic trough, so that the extremity of the tube comes under the gas jar. Bubbles of gas rise to the top of the jar, and gradually displace the water. Other jars should be standing on the shelf of the trough, ready filled with water, to take the place of the one filled with gas.

2. Jet of inflamed hydrogen.

3. Arrangement for filling bladders at the pneumatic trough.

4. Soap bubbles blown from a bladder filled with hydrogen, or a mixture of oxygen and hydrogen.

5.



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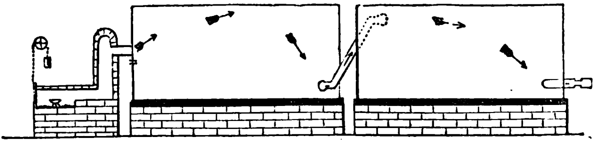


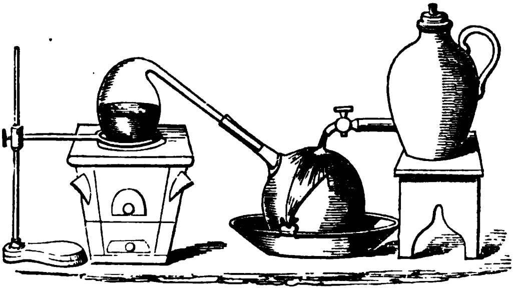
5. Arrangement for preparing oxygen and gases requiring heat.

6. Iron wire burnt in oxygen.

7. Preparation of nitrogen by burning phosphorus in a confined portion of air. The gas jar stands in a deep dish of water.

8. Section of the arrangement for the manufacture of sulphuric acid from burning sulphur and the nitrate of potash.

9.



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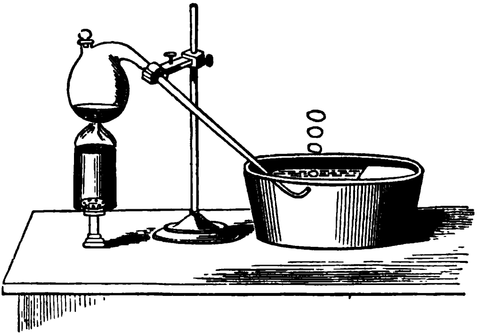
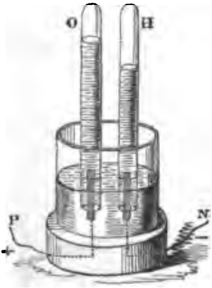


9. Arrangement for distillation, or the preparation of nitric acid, from the nitrate of potash and sulphuric acid.

10. Phosphuretted hydrogen apparatus.

11.



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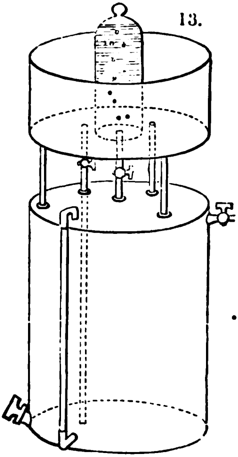

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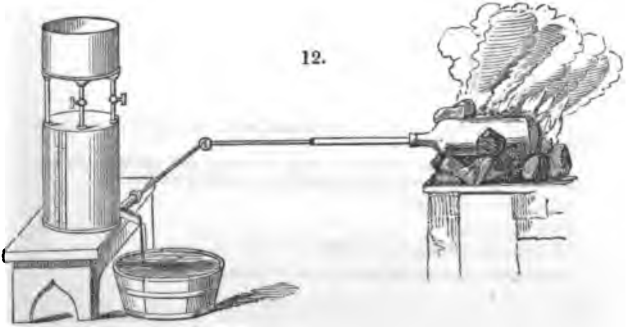
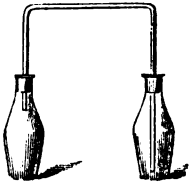


11. Decomposition of water by a galvanic battery.
12. Preparation of oxygen from the black oxide of manganese.
13. Gas holder, used when a large quantity of gas is required.

14.



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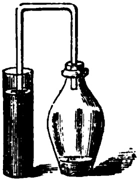

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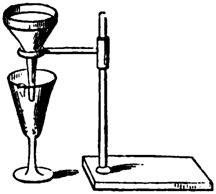

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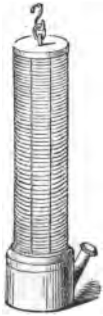

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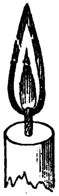

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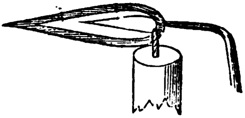

14. Apparatus for preparing carbonic acid, and gases heavier than common air, and not requiring heat.

15. Arrangement showing the passing of sulphuretted hydrogen through a metallic solution, or for dissolving the gas in distilled water.

16. Filtering apparatus.

17. Davy Lamp.

18. Diagram showing the structure of flame.

19. Diagram showing the structure of the blowpipe flame.

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